STUDENT'S ESSENTIALS

OF

CHEMISTRY,

THEORETICAL AND INORGANIC,

BY

CONDICT W. CUTLER, M.D.,

SURGEON TO EASTERN DISPENSARY, NEW YORK,
PHYSICIAN TO NEW YORK DISPENSARY, &C.

SURGEON GEN'L'S OFFICE
LIBRARY

NEW YORK,
1885.

Entered according to act of Congress, in the year 1885,
BY CONDICT W. CUTLER,
In the Office of the Librarian of Congress, at Washington.

"BANNER" STEAM PRINT,
Morristown, N. J.

PREFACE.

——:o:——

In writing this little book I have endeavored to clearly set before the student such essentials of theoretical and inorganic chemistry as will, on the one hand, give him a thorough fundamental knowledge of the subject; and on the other, to enable him to better understand and appreciate the more advanced works on chemistry.

150 *W.* 53*rd Street, New York.*
July, 1885.

CONTENTS.

—:o:—

PART I.—ON THEORETICAL CHEMISTRY.

PART II.—ON INORGANIC CHEMISTRY.

Chapter.

I. HYDROGEN.
II. OXYGEN.
III. NITROGEN.
IV. CARBON.
V. COMBUSTION.
VI. HALOGENS.
VII. SULPHUR.
VIII. SILICON AND BORON.
IX. PHOSPHORUS.
X. ARSENIC.
IX. ANTIMONY AND BISMUTH.
XII. TIN, ZINC AND CADMIUM.
XIII. CHROMIUM, MANGANESE AND ALUMINUM.

XIV. Iron, Cobalt and Nickel.
XV. Platinum, Gold and Silver.
XVI. Lead, Mercury and Copper.
XVII. Magnesium, Strontium, Barium and Calcium.
XVIII. Lithium, Potassium, Sodium and Ammonium.
XIX. Experimental Chemistry.

PART I.

THEORETICAL CHEMISTRY.

THEORETICAL CHEMISTRY.

——:o:——

Chemistry treats of the atomic composition of matter, and of those changes in matter, which result from a re-arrangement of these atoms.

Matter is anything we can appreciate by our senses, and is made up of *sixty-seven elements* or substances that can not be divided into other substances.

Matter may be divided into :—

Atoms.

Molecules.

Masses.

Atoms are the smallest divisions of matter and unite to form molecules.

A *Molecule* is the smallest group of atoms, that will exhibit the physical and chemical properties of the substance, and

can not be divided without losing these characteristics.

A *Mass* is a collection of molecules.

COMPOSITION OF MOLECULES.

The composition of a molecule is ascertained by

a. *Analysis*, which consists in separating the molecule into its constituent atoms; and,

b. *Synthesis*, which consists in uniting the constituent atoms to form the molecule.

The atoms thus composing any molecule may be alike, forming *elemental* molecules; or unlike, forming *compound* molecules. Rearrangement of the atoms in any elemental molecule yields no new kind of matter as does the rearrangement of atoms in compound molecules.

There have been found but sixty-seven such elemental molecules or substances, and these have received the name of *elements*.

From these sixty-seven kinds of atoms, all masses of matter are made up.

These elements are all named, and each

name has, for convenience, received a symbol. They are as follows :—

TABLE OF ELEMENTARY BODIES, WITH THEIR SYMBOLS AND ATOMIC WEIGHTS.

TABLE I.

Name.	Symbol.	Atomic Weight.
Aluminum,	Al	27.4
Antimony,	Sb	122
Arsenic,	As	75
Barium,	Ba	137
Beryllium,	Be	9.4
Bismuth,	Bi	210
Boron,	B	11
Bromine,	Br	80
Cadmium,	Cd	112
Caesium,	Cs	133
Calcium,	Ca	40
Carbon,	C	12
Cerium,	Ce	92
Chlorine,	Cl	35.5
Chromium,	Cr	52.2
Cobalt,	Co	58.8
Copper,	Cu	63.4
Didymium,	D	95
Erbium,	E	168.9
Fluorene,	F	19
Gallium,	Ga	68

Name.	Symbol.	Atomic Weight.
Gold,	Au	197
Hydrogen,	H	1
Indium,	In	113.4
Iodine,	I	127
Iridium,	Ir	198
Iron,	Fe	56
Lanthanum,	La	93.6
Lead,	Pb	207
Lithium,	Li	7
Magnesium,	Mg	24
Manganese,	Mn	55
Mercury,	Hg	200
Molybdenum,	Mo	96
Nickel,	Ni	58.8
Niobium,	Nb	94
Nitrogen,	N	14
Osmium,	Os	199.2
Oxygen,	O	16
Palladium,	Pd	106.6
Phosphorus,	P	31
Platinum,	Pt	197.4
Potassium,	K	39.1
Rhodium,	Rh	104.4
Rubidium,	Rb	85.4
Ruthenium,	Ru	104.4
Selenium,	Se	79.4

Name.	Symbol.	Atomic Weight.
Silicium,	Si	28
Silver,	Ag	108
Sodium,	Ma	23
Strontium,	Sr	87.6
Sulphur,	S	32
Tantalum,	Ta	182
Tellurium,	Te	128
Thalium,	Tl	204
Thorinum,	Th	235
Tin,	Sn	118
Titanium,	Ti	50
Tungsten,	Ui	184
Uranium,	U	240
Vanadium,	V	51.2
Yttrium,	Y	92
Zinc,	Zn	65.2
Zirconium,	Zr	89.6

NUMBER OF ATOMS IN ELEMENTAL MOLECULES.

From the law of Ampere - *Equal volumes of all bodies in the gaseous state contain the same number of molecules*, it follows, *that the molecules of all bodies when in the gaseous state, are of the same size; and their weight compared to that of hydrogen is proportional to the weight of any given volume, compared*

with the same volume of hydrogen. For example, if one volume of hydrogen gas contains one hundred molecules, then an equal volume of chlorine will contain one hundred molecules.

If these volumes be mixed together, and exposed to the sun light, they will unite and form two volumes of hydrochloric acid-gas, containing two hundred molecules. Upon analysis, each molecule of this gas is found to contain two atoms, one of hydrogen and the other of chlorine, hence the two hundred molecules of this gas will contain two hundred hydrogen atoms, and two hundred chlorine atoms.

As each two hundred hydrogen and chlorine atoms came from one hundred molecules, then each molecule of hydrogen and chlorine must contain two atoms.

In this way it has been found that most of the elements contain two atoms to the molecule, but not all.

The *atomicity* of any molecule is the number of atoms which it contains. The fol-

lowing table gives the atomicity some of the more important of the elements:

TABLE II.

Mon-Atomic.	Di-Atomic.
Mercury,	Hydrogen,
Cadmium,	Oxygen,
Zinc,	Chlorine,
Barium,	Iodine,
	Bromide.
	Potassium,
	Nitrogen,
	Sulphur, &c.,
	Tetre-Atomic.
	Phosphorus,
	Arsenic,
	Hex-Atomic.
Tri-Atomic.	Sulphur.
Oxygen,	
(as ozone).	

ATOMIC AND MOLECULAR WEIGHTS.

The *Atomic weight* of any atom is its relative weight compared to hydrogen which is taken as unity (1).

The *Molecular weight* of any molecule is the sum of the atomic weight of its constituents.

To determine the atomic weight of any atom we must know—

a. The quantity by weight of the substance combining with one atom of hydrogen, and—

b. The molecular weight of the hydrogen compound. *For example.* By analysis we know that in one hundred parts of hydrochloric acid gas there are 97. + parts of chlorine and 2. + parts of hydrogen.

The proportion 2. + : 97. + : : 1 : (35. +) represents the quantity by weight of chlorine combining with 1 part of hydrogen, which is 35.4. The density of any gas is one-half its molecular weight, hence if the density of hydrochloric acid gas is 18. + its molecular weight must be 36. +.

But we have already found that the combining weight of hydrogen and chlorine was 1 and 35. + or 36. +. Therefore as the sum of the combining and molecular weights are the same, the gas must contain an equal number of hydrogen and chlorine atoms. As the atomic weight of hydrogen is taken as 1, that of chlorine will be 36. + — 1 or 35. +.

If on the contrary the combining weight

was only one-half or one-third the molecular weight, then a molecule of the gas would contain two or three times as much chlorine and hydrogen and the atomic weight of the chlorine would be 35. + x 2 or 35. + x 3.

In some cases, however, an element does not combine directly with hydrogen, as is the case with silver. Then a comparison is made with some other element, as chlorine, which has a constant combining power with hydrogen of 35. +, and in this way its atomic weight is ascertained. *For example.* Chloride of silver yields on analysis 75. + per cent. silver and 24. + per cent. chlorine. Then the proportion 24 + : 75 + : : 35. + : (108), will give the atomic weight of silver as 108. (See table 1, for atomic weights.)

ATOMIC EQUIVALENCE.

The *equivalence* of any atom is the quantity of its combining power expressed in hydrogen units, which has a combining power of 1.

These atoms having the combining power of one (1) are called *monads*—of two (2) *diads*—of three (3) *triads*, &c. Many atoms

have more than one combining power, that is. they may have an equivalence of 1, 3, 5, &c., or 2, 4, 6, &c., but this combining power in the same kind of atoms always increases or diminishes by two.

Atoms whose equivalence is even are called *artiads;* those whose equivalence is odd, *perissads.* The following table gives the most constant equivalence of elements:

TABLE III.

Monads.	Diads.
Hydrogen,	Strontium.
Fluorine,	Barium,
Chlorine.	Magnesium,
Bromine,	Zinc,
Iodine.	Cadmium,
Lithium.	Cerium,
Potassium.	Mercury,
Silver.	Copper,
Triads.	Oxygen,
Antimony,	Sulphur,
Bismuth.	Calcium.
Boron,	Tetrads.
Gold,	Carbon,
Nitrogen,	Silicon,
Phosphorus.	Tin,
Arsenic.	Aluminum,

Hexads.	Tetrads.
Chrominum,	Platinum,
Manganese,	Lead.
Iron,	
Cobalt,	
Nickel.	

For example. Hydrogen is a monad and chlorine is a monad, consequently one atom of hydrogen will just saturate one atom of chlorine. Again, hydrogen being a monad and oxygen a diad, it will take two atoms of hydrogen to saturate one of oxygen, &c.

The *graphic symbol* of any atom is one indicated by a circle with lines radiating from it equal in number to its combining power. Thus an atom of hydrogen would be indicated by a circle with one radiating line—oxygen with two lines, &c.

A few of the elements, the most important of which are :—

Mercury,
Copper,
Iron,
Cobalt and
Nickel,

besides *one* atom, with an equivalence of 2,

4, 6, &c., uniting with other atoms, they frequently combine by the use of *two* atoms. Thus one atom of mercury (Hg), which is a diad, will combine with two of chlorine which is a monad or two atoms of mercury (Hg_2), will combine with two of chlorine thus making entirely different compounds. ($Hg\ Cl_2$.) and ($Hg_2\ Cl_2$.)

MULTIPLICATION OF ATOMS AND MOLECULES.

Atoms are multiplied by placing a numeral below and to the right of its symbol. Thus C indicates one atom of carbon—C_2. two atoms of carbon, &c.

Molecules are multiplied in the same way after inclosing their symbols in brackets, or by placing the numerals before the molecular formula Thus, $(H_2O)_2$ or $2H_2O$ would represent two molecules of water—$(H_2O)_3$, or $3H_2O$ three molecules of water, &c.

When the number of atoms of any constituent in a molecule is more than one, this is indicated by a numeral placed below and to the right of its symbol. Thus CO_2 would represent one molecule of carbon dioxide.

CHEMICAL FORMULAS.

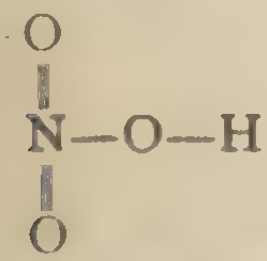
A *compound molecule* is one whose constituent atoms are unlike.

The representation of any compound molecule by symbols is called a *formula*. Thus, HCl is hydrochloric acid, H_2O, water, &c.

Formulas may be empirical or rational.

An *empirical* formula is one derived from analysis, and is merely an expression of the kind and relative, or absolute number of atoms in any molecule.

A *rational* formula expresses the kind and absolute number of atoms, and also how these atoms are arranged. Thus, HNO_3 is the empirical formula for nitric acid, while



is its rational formula. H_2O the empirical formula and H—O—H the rational formula for water.

NAMING MOLECULES.

The names of all compound molecules

are derived from those of their constituents.

When the constituent atoms have but one equivalence as sodium and chlorine, the termination *ide* is generally added. Thus, NaCl is sodium chlor*ide*. Magnesium and oxygen each have an equivalence of two, hence Mg=O is magnesium ox*ide*, &c.

In cases, however, where the constituent atoms have more than one equivalence, as for example nitrogen which has an equivalence of 3 and 5 usually, then the compound with the lower equivalent takes the termination *ous*, and the higher one *ic*. Thus nitrogen a triad uniting with oxygen a diad, the result is nitr*ous* oxide. Again, nitrogen a pentad uniting with oxygen a diad the result is nitr*ic* oxide.

Compound molecules having a lower equivalence still than the ous, the prefix *hypo* is attached, or if higher than the ic, the prefix *per*. Thus, theoretically, nitrogen may be a monad or a septad, and if united to oxygen a diad, there would be formed *hypo*-nitrous oxide and *per*-nitric oxide.

In a few cases the number of atoms com-

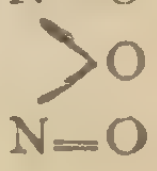
bining to form the molecule is indicated in naming the substance. Thus, CO_2 is called carbon di-oxide, because there are two atoms of oxygen which distinguishes it from CO, carbon, mon-oxide.

CONSTRUCTION OF MOLECULES.

When atoms with different equivalents unite to form compounds, each atom must furnish the same number of bonds, this number being the least common multiple of their atomic equivalence; and the number of atoms required by each constituent is obtained by dividing the least common multiple by the equivalence of each atom.

Thus, if nitrogen is a triad and oxygen a diad, both must furnish six bonds ($2 \times 3 = 6$), the least common multiple of 2 and 3. To furnish these bonds, nitrogen which already has 3 bonds must furnish two atoms ($6 \div 3 = 2$), while oxygen having but two bonds, must furnish three atoms ($6 \div 2 = 3$).

The formula formed by this union of nitrogen and oxygen must be N_2O_3,



SATURATED AND UNSATURATED MOLECULES.

When the bonds of all the atoms constituting a molecule are mutually combined the molecule is called *saturated*, but when some of the bonds remain free the molecule is *unsaturated*, and the group of atoms composing it form a *compound* radical.

When the unsaturated atoms are all of the same kind, the group is called a *simple* radical.

As radicals cannot exist in a free state in nature, they enter into combination with other atoms as do single atoms, and take the termination *yl*. Thus, in water (H_2O,) if one hydrogen is removed there remains (HO), a compound radical called hydrox*yl*.

As hydrogen is a monad and oxygen a diad, (HO) has one free bond which may unite with any atom having one free bond, as potassium, forming KOH, or with other radicals having one free bond as (NO), thus forming HNO_2.

In some cases artiad radicals, or compound radicals with an even number of free bonds, can exist free in nature by mu-

tual saturation of these free bonds. Thus CO, carbonal or carbon-mon-oxide is a diad radical, carbon having four bonds, only two of which are saturated by oxygen, but as the remaining bonds saturate each other (<C=O) it may exist free in nature.

In perissad radicals, however, this can never take place, for there always exists an odd number of free bonds, consequently the union of any one free bond with another would still leave the molecule unsaturated.

It is by this mutual saturation of bonds that a triad may become a monad or a tetrad a diad, &c.

NEGATIVE AND POSITIVE ATOMS.

In considering the formation of acids, bases and salts, the atoms constituting them are either positive or negative.

A *positive atom* is one which is attracted to the negative pole in electrolysis.

A *negative atom* is one which is attracted to the positive pole in electrolysis.

The following table gives the most im-

portant of the positive and negative elements :

TABLE IV.

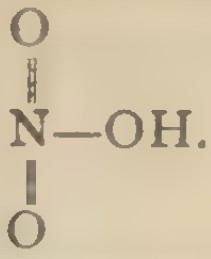
Negative.	Positive.
Oxygen,	Hydrogen,
Nitrogen,	Platinum,
Carbon,	Aluminum,
Bromine,	Gold,
Iodine,	Copper,
Fluorine,	Ammonium,
Phosphorus,	Calcium,
Sulphur,	Potassium,
Arsenic,	Lithium,
Antimony,	And the metallic
Boron,	elements gener-
Bismuth,	ally.
Silicon.	

THE FORMATION OF ACIDS, BASES AND SALTS.

An *acid molecule* is one consisting of one or more negative atoms united by oxygen to hydrogen, and has the property of turning vegetable blues - as litmus, red.

Thus in nitric acid, (HNO_3,) nitrogen is a

negative element united by oxygen to hydrogen,



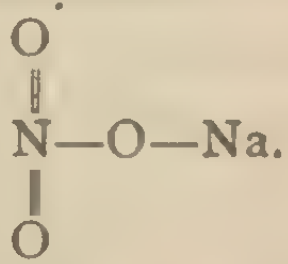
A *basic molecule* is one consisting of one or more positive atoms united by oxygen to hydrogen.

Thus, sodium hydrate, (Na OH,) is a base. Sodium is a positive atom united by oxygen to hydrogen, Na—O—H.

Many chemists define a base as a metallic oxide, and a basic element as a metallic element.

A *saline molecule* is one consisting of one or more positive atoms united by oxygen to one or more negative atoms.

Thus nitrate of soda, ($NaNO_3$) is a salt. Sodium is a positive atom united by oxygen to nitrogen, a negative atom,



These acids, bases and salts are formed on the plan of the structure of water, (H—O—H,)

commonly called the *water type*. Thus exchanging for one of the hydrogens a negative atom an acid is formed, for a positive atom a base—and by exchanging both hydrogens, one for a negative and the other for a positive atom, a salt is produced.

Chemically speaking, a salt is usually produced by the action of an acid upon a base, or, theoretically, by replacing one or more of the hydrogens of any acid by basic elements.

There are three kinds of salts, called acid, normal and double. An *acid salt* is one in which all the hydrogens of the acid have not been replaced by basic elements, as disodium phosphate ($Na_2 HPO_4$).

An acid salt sometimes takes the prefix of *super* or *bi* as in *super*-phosphates or *bi*-carbonates.

A *normal salt* is one in which all the hydrogens have been replaced, as sodium phosphate ($Na_3 PO_4$).

A *double salt* is one in which the acid hydrogens are replaced by more than one kind of basic elements, as ammonio-magnesium phosphate (NH_4), Mg PO_4.

NAMING ACIDS AND SALTS.

As compound molecules take the terminations ous and ic and with prefixes, hypo and per, according to the number of equivalence of their constituent atoms, so when acids are formed by the union of hydrogen by oxygen to these compound molecules they also take like terminations.

Thus hyponitrous oxide united to hydrogen by oxygen forms hyponitrous acid. In the same way is formed nitrous acid and nitric acid.

In a like manner any salt produced by the action of an *ous* acid forms an *ite* salt, while an *ic* acid forms an *ate* salt either with or without the prefixes hypo and per. Thus we have by the action of nitr*ous* acid on sodium hydrate the formation of sodium nitr*ite*, while the action of nitr*ic* acid on the same base forms sodium nitr*ate*.

TYPES FOR THE FORMATION OF COMPOUNDS.

Beside the water type of compounds there are several other types, the most important being the *hydrochloric acid type*, the *ammonia type*, and the *marsh gas type*.

In the *hydrochloric acid type* H—Cl. by

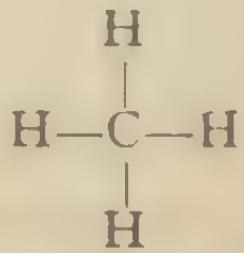
the replacing of the hydrogen atom by other atoms, as sodium, potassium, &c., a new class of compounds are formed, as sodium and potassium chloride, &c.

The *ammonia type*

is also an important type, especially in the formation of organic compounds. These compounds are formed by replacing one or more of the hydrogens by other atoms. Thus by replacing one hydrogen there remains (NH_2), called *amodogen*, which acting as a monad radical can unite with the monad potassium forming (NH_2) K, or with the diad calcium, by taking two molecules of the radical, forming $(NH_2)_2$ Ca.

If a hydrogen atom in the ammonia type is replaced by a positive atom it forms an *amine*—by a negative atom, an *amide*—by both a negative and a positive atom, an *alkalamine*.

The *marsh gas type*,



is also an important type in organic chem-

istry. New compounds are formed by substituting other atoms in the place of one or more hydrogens.

CHEMICAL REACTIONS.

All molecules are subject to chemical change, and any change taking place among the atoms constituting any molecule is called *chemical reaction.*

This change is expressed by molecular formulæ, the chemical reactions being represented in the form of equations. Thus let AB be one molecule of any substance to be acted upon by another molecule CD, then the equation would be expressed thus : AB $+$CD$=$AD$+$CB. As no loss of weight can be the result of any chemical change, then the molecular weight of the two sides of the equation must be equal.

If the chemical change results in the production of any substance insoluble in the menstruum employed, such a compound will appear as a precipitate. If, on the contrary, a volatile substance is produced by such a chemical change it will be evolved as a gas or vapor.

DIFFUSION OF GASES.

By diffusion of gases is meant the passage of one or more gases into another.

The velocity of diffusion of any gas is inversely as the square root of its density.

PART II.

INORGANIC CHEMISTRY.

—:o:—

CHAPTER I.

Inorganic chemistry treats of all the elements both metallic and non-metallic in their free and combined states as they occur in the inorganic world.

The chemistry of plants and animals is treated of under organic chemistry, or chemistry of the carbon compounds.

In the consideration of each element, the following order of arrangement will assist the student in memorizing the facts in their most important relations :—

I. The name of the element, its symbol, atomic weight and equivalence.

II. Its history.

III. Its occurrence.

IV. Its preparation.

V. Its properties.

a. Physical,

b. Chemical.

VI. Its uses.

VII. Its compounds.

HYDROGEN.

Symbol H. *Atomic weight* I. *Equivalence* I.

History. Hydrogen was discovered in 1766 by Cavendish, who described it under the name of inflammable air.

Occurrence. Hydrogen exists free in certain volcanic gases, in or around the sun and some of the fixed stars.

In the combined state it is found in water, in all animal and vegetable tissues, petroleum, &c.

Preparation. Hydrogen can be prepared by both physical and chemical forces. As an example of its preparation by *physical* agents, water (H_2O) by means of electrolysis will separate into its constituent atoms, oxygen collecting at the positive and hydrogen at the negative pole, thus :—

Water. Hydrogen. Oxygen.

$$H_2O = H_2 + O.$$

Chemically, hydrogen is prepared in a number of ways. For example if metallic sodium is thrown upon water, the water is

decomposed, and hydrogen is set free. Thus :—

Water. Sodium. Sodium hydrate. Hydrogen.

$$H_2O + Na = NaOH + H.$$

Again, it may be prepared by the action of sulphuric acid upon metallic zinc, the sulphuric acid being decomposed and hydrogen set free. Thus :—

Sulphuric acid. Zinc. Zinc sulphate, Hydrogen.

$$H_2So_4 + Zn = ZnSo_4 + H_2.$$

Properties.

a. *Physical.* Hydrogen is a colorless, odorless, and tasteless gas. It is the lightest form of matter known, hence its diffusibility is greater than any other gas. It is but slightly soluble in water.

b. *Chemical.* Hydrogen is combustible—that is, at an elevated temperature it will combine with oxygen with the evolution of light and heat. It will not support combustion. It burns with a pale blue flame, and with an intense heat. Its product of combustion is water.

Uses. Hydrogen on account of its lightness is used to inflate balloons. It is used also as a heating agent on account of its high temperature produced by combustion.

CHAPTER II.

OXYGEN.

Symbol O. *Atomic weight* 16. *Equivalence* II.

History. Oxygen was first discovered by Priestly in 1774, and described by him as dephlogisticated air. Sheele discovered it independently the next year, and called it empirical air.

Occurrence. Oxygen is the most abundant element in nature. It forms one-fifth part of the atomosphere where it exists free. In the combined state it forms two-thirds of the globe—water being two-thirds and minerals one-half oxygen by weight. It is found combined with other elements in all animal and vegetable tissues.

Preparation. Oxygen may also be prepared by means of physical and chemical forces. By the aid of *physical* agents, oxygen can be prepared by *heat, light* and *electricity.* Thus for example of its prepara-

tion by heat,—Mercuric oxide exposed to a high temperature is decomposed and metallic mercury and oxygen are produced. Thus:—

Mercuric-oxide. Mercury. Oxygen.

$$HgO = Hg + O.$$

or by light,—Carbon di-oxide (CO_2) in the leaves of plants is decomposed by the sunlight into carbon and oxygen. Thus:—

Carbon-dioxide. Carbon. Oxygen.

$$CO_2 = C + O_2.$$

or by electricity as seen in the electrolysis of water described under the chapter on hydrogen.

Chemically, oxygen is prepared in a number of ways. It is prepared from the oxygen of the air in the following manner:

1. Pure air is passed over a heated mixture of manganese di-oxide (MnO_2) and sodium hydrate ($NaOH$), thus forming sodium manganate (Na_2MnO_4) and water H_2O. Thus:

Mang. dioxide. Sod. hydrate. Sod. manganate. Water.

$$MnO_2 + NaOH + air = Na_2MnO_4 + H_2O.$$

2. This heated sodium manganate (Na_2MnO_4) is then raised to a high temperature, and a current of steam passed over it which

decomposes it into its original constituents again, and setting oxygen free from its mixture with nitrogen. Thus :—

Sod. manganate. Water. Mang dioxide. Sod. hydrate. Oxygen.

$$Na_2MnO_4 + H_2O = MnO_2 + 2NaOH + O.$$

This process can be kept up continually, and oxygen manufactured in large quantities.

The usual method of preparing oxygen for commerce is by heating chlorate of potash ($KClO_3$), which yields a large quantity of oxygen. Thus :—

Potass. chlorate. Potass. chloride. Oxygen.

$$KClO_3 = KCl + 3O.$$

Properties.

a. *Physical.* Oxygen is a colorless, odorless and tasteless gas. It is a little heavier than air and but slightly soluble in water.

b. *Chemical.* Oxygen enters into combination with all the elements except fluorine.

The combination of any element with oxygen really constitutes combustion, but the term is usually restricted to those combinations of oxygen which are accompanied with heat and light, the usual result if the oxidation is a rapid one.

As the result of oxidation, oxides are produced.

Uses. Oxygen is used for aiding combustion for purposes of either heat or light.

In the air diluted with nitrogen its uses for respiratory purposes are well known.

COMPOUNDS.

COMPOUNDS OF OXYGEN AND HYDROGEN.

HYDROGEN OXIDE OR WATER.

Formula, H_2O or H—O—H.

History. Water was considered an elementary substance until Lavoinier in 1773, proved its compound nature.

Occurrence. Water occurs abundantly in nature, both free and combined. Occurring in the free state it is either—

Atmospheric or

Terrestrial.

I. Atmospheric water occurs in the form of rain, snow, hail, mist, &c.

II. Terrestrial waters are—

Sweet,

Salt or

Mineral.

a. *Sweet* water is found in—

1. Springs.

2. Rivers.
3. Lakes and
4. Wells.

b. *Salt* water is found in—

1. Oceans and
2. Inland seas.

c. *Mineral* waters are those that contain an excess of mineral ingredients, and are named from their principal constituents. Thus :—

1. Saline.
2. Alkaline.
3. Chalybeate.
4. Acid—containing free acids.
5. Alum.
6. Acidulous—containing carbonic acid.
7. Sulphur.
8. Borax.
9. Silicious.

Water in the combined state presents four degrees of combination as—

a. Hydroscopic moisture, or that which can be driven off at a temperature of 212° F.

b. Water of crystallization.

c. Chemical combination, also called hydration,

d. Solvent. Not always a mere mechanical solution for heat or cold is often the result.

Ingredients of water. Natural waters are seldom pure, that is they contain more than the constituent elements, hydrogen and oxygen.

Water always contains a certain amount of mineral matter, such as potassium, sodium, calcium and magnesium in the form of chlorides, bicarbonates and sulphates. Beside this mineral matter it also contains organic matter, derived from the decomposition of plants, and gases in solution derived from the air.

Preparation. Water is always the product of combustion in presence of oxygen. It can be prepared by direct union of its constituents mixed in proper proportion by means of an electric spark. It is also the result of many chemical combinations.

Properties.

a. *Physical.* Water is a colorless, odorless and tasteless liquid. A poor conductor of heat and electricity. It is taken as the standard for the specific gravity of liquids. Its maximum density is at 4°C, it expanding

for all lower and higher temperatures. It freezes at o°C, forming a crystalline solid —ice ; and is converted into vapor or steam at 100°C.

Most gases are soluble in water, it dissolves them in proportion to their *volume* and not their density.

b. *Chemically*, water is very active, entering into chemical combination with many substances. Thus :—

Calcic oxide. Water. Calcic hydrate.

$$CaO + H_2O = Ca(OH)_2.$$

Efflorescent substances are those that lose their water of crystallization and become powdered.

Deliquescent substances are those that attract moisture and liquify.

Uses. Water is used for manufacturing and domestic purposes, and may contain ingredients which make it unfit for either.

The ingredients present which render the water harmful for manufacturing purposes are the minerals, especially the calcium salts which are deposited in the boilers as calcium carbonate ($CaCO_3$). This salt is not soluble in pure water, but is soluble in

water containing carbonic acid (H_2CO_3). Thus:—

Calc. carbonate. Carbonic acid. Calc. bi-carbonate.

$$CaCO_3 + H_2CO_3 = CaH_2(CO_3)_2.$$

This salt is soluble in water, but upon the application of heat it is decomposed into carbonic dioxide and calcium carbonate. Thus:—

Calc. bi-carbonate. Water. Carbon dioxide. Calc. carbonate

$$CaH_2(CO_3)_2 = H_2O + CO_2 + CaCO_3.$$

This calcium carbonate depositing on the inside of boilers destroys them. An apparatus called the calcium catcher is sometimes attached to boilers to remedy this danger. Saw-dust and other foreign substances is sometimes put in the boilers for the calcium to deposit upon instead of upon their iron plates.

In water for domestic use the presence of mineral matter does no harm unless the water contains much calcium which renders it "hard," and hence poor for washing and cooking purposes.

It is here, however, that the organic matter present may or may not prove very harmful to health. If the organic matter present in the water be due to the decompo-

sition of vegetable matter, and it is not great in amount, it will do no harm.

If, on the contrary, the organic matter is the result of animal impurities, sewage, &c., in which disease germs are often present, it may be the means of causing much sickness. Boiling the water will destroy these impurities.

The presence of much chlorine or ammonia in the water renders it suspicious—chlorine, for it indicates sewage contamination, and ammonia, organic matter.

Tests.

a. For *chlorine.* Add to the water a little nitric acid (HNO_3) and then nitrate of silver ($AgNO_3$). The chlorine is precipitated in white flakes, as the chloride of silver.

b. For *organic matter.* Add permanganate of potash to the water; if much organic matter is present it will oxidize the potash solution and thus destroy its color.

c. For *ammonia.* Add to the water a little "Nessler reagent,"—a solution of iodide of mercury (HgI_2), and caustic potash (KOH), and if ammonia is present it will turn the water a dark yellow color; more the ammonia, darker the solution.

HYDROGEN-DIOXIDE OR FREE HYDROXYL.

Formula. H_2O_2. H—O—O—H.

Preparation. It is always the result of slow oxidation and is prepared by the action of carbonic or hydrochloric acid (H_2CO_3 or HCl) upon barium oxide. Allowing the water present to evaporate slowly, leaves the H_2O_2 behind as a syrupy liquid. Thus :

Barium oxide. Carbonic acid. Barium carbon, Free hydroxyl.

$$BaO_2 + H_2CO_3 = BaCO_3 + H_2O_2.$$

Properties. Hydrogen dioxide is a colorless syrupy liquid which decomposes at temperatures over 100° into water and oxygen. As it evolves oxygen readily, it is an oxidizing agent.

Uses. It is used in bleaching vegetable colors.

OZONE.

Symbol O_3. *Molecular weight* 48.

Ozone is not an element but a form of oxygen with three instead of two atoms to a molecule, and sometimes described as active oxygen.

History. Although ozone was really discovered in the eighteenth century, Schonbein in 1840 was the first to describe it.

Occurrence. Ozone is found free in the

air, especially after a thunder shower. It is probably produced also in the atmosphere by the growth of plants.

Preparation. Ozone can be prepared by chemical and physical forces. Electricity is the best *physical* agent for the production of ozone. By its passage through the air, as by lightning, or by the electrolysis of water acidulated with hydrochloric acid (HCl), ozone is formed.

Chemically, ozone is prepared by the slow oxidation of phosphorus or sulphur. Thus:

Phosphorus. Oxygen. Phosphoric oxide. Ozone.

$$P_2 + 3O_2 = P_2O_3 + O_3.$$

or by the rapid oxidation of red hot platinum.

Properties.

a. *Physical.* Ozone is a colorless gas with an odor like that of weak chlorine. It is a little heavier than oxygen.

b. *Chemically*, it is a more active oxidizer than oxygen.

It bleaches strongly and is poisonous to animal life, although easily converted into oxygen. It decomposes potassium iodide (KI), setting free iodine vapor which colors

starch solution blue. This acts as a test for the presence of ozone.

Uses. Ozone is a powerful bleaching agent, and as it readily oxidizes it is used by many as a disinfectant for water closets, &c., to destroy foul odors.

CHAPTER III.

NITROGEN.

Symbol N. *Atomic weight* 14. *Equivalence* I–III–V.

History. Nitrogen was discovered by Rutherford, in 1772, in the air which had been breathed by animals.

Occurrence. Nitrogen is found free in the atmosphere. In the combined state it is found in the ammonia compounds, nitrates and nitrites, and in many of the animal and vegetable tissues.

Preparation. The easiest way to obtain nitrogen is from the air where it exists in a free state. This is accomplished by using the accompanying oxygen by combustion—in other words by burning out the oxygen.

Chemically it is prepared by heating ammonium nitrite (NH_4NO_2), thus breaking it up into water and nitrogen. Thus:—

Ammonium nitrite Water. Nitrogen.

$$(NH_4)NO_2 = (H_2O)_2 + N_2.$$

Properties.

a. *Physical.* Nitrogen is a colorless, odorless and tasteless gas. It is lighter than air and soluble in water.

b. *Chemical.* Nitrogen is very inert, entering directly into combination with very few elements. In combination, however, it is very active, forming corrosive and explosive mixtures as nitric acid, nitro-glycerine, &c. It will not support combustion.

Uses. One of the principal uses for nitrogen is a natural one in diluting the oxygen of the air so as to render it suitable for respiratory purposes, and for limiting its oxidizing properties.

THE ATMOSPHERE.

Composition. The atmosphere is made up of a mixture of normal and abnormal ingredients. These are:—

I. *Normal.*	Parts per 100.
1. Oxogen,	20.50
2. Nitrogen,	78.00
3. Watery vapor,	1.40
4. Carbon dioxide,	.05
5. Ozone,	trace
6. Gaseous hydrocarbons,	trace

7.	Hydrogen sulphide,	trace
8.	Ammonia,	trace
9.	Acids,	trace
	Total,	100.00

II. *Abnormal.*

A. Dead matter.

a. Mineral.

1. Sand.
2. Clay.
3. Cosmical dust.
4. Sodium chloride. } From sea water.
5. Calcium sulphate. }

b. Non-mineral.

1. Animal substances.
2. Vegetable substances,
3. Excrements.
4. Products of decomposition.
5. Sewer gas.

B. Live matter.

1. Pollen of plants.
2. Germs. { Disease germs. / Fungi, / Yeast plant.

Properties.

a. *Physical.* Air is one of the standards used for the specific gravity of gases. It exerts a pressure of 15 lbs. to the square

inch on the surface of the earth. The height of the atmosphere is not known, but probably exists for fifty miles above the earth, its density diminishing as it ascends.

b. *Chemical.* It is a mixture of oxygen and nitrogen and not a chemical union. The gases do not separate on account of the law governing their diffusion.

The atmosphere is a good supporter of combustion and a poor conductor of heat and electricity. It is slightly soluble in water.

Uses. The most important use of the atmosphere is for respiratory purposes.

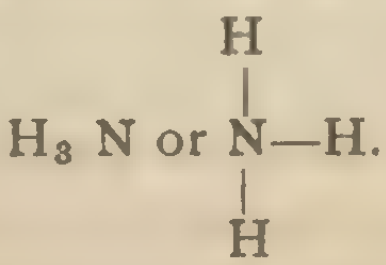
The most deleterious ingredient in the air is the disease germs. These are especially found in sewer gas and are conveyed to the air mechanically from sewage by bubbles.

COMPOUNDS.

a. OF NITROGEN AND HYDROGEN.

HYDROGEN NITRIDE, OR AMMONIA.

Formula.



History. Ammonia has been known for

many centuries, but the name was not given it until 1781 by Bregman.

Occurrence. It occurs but sparingly in nature, a little being found in water, air, soil and animal matter.

Preparation. The compounds of ammonia are usually prepared by the destructive distillation of animal matter or from the ammoniacal liquor obtained from distillation of coal.

Ammonia gas is best prepared by the action of quick lime (CaO) upon ammonium chloride (NH_4Cl). Thus:

Ammonium chloride. Calcic oxide. Calcic chloride. Water.

$$(NH_4Cl)_2 + CaO = CaCl_2 + H_2O$$

Ammonia.

$+(H_3N)_2$.

a. *Physical.* Ammonia is a colorless, pungent and alkaline gas, turning red litmus paper blue. It is very volatile, lighter than air and freely soluble in water.

b. *Chemically*, it will not support combustion and is slightly combustible at elevated temperatures. It is readily decomposed by heat and electricity into nitrogen and water.

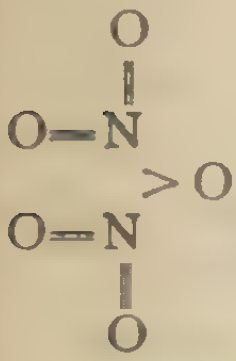
b. OF NITROGEN AND OXYGEN.

As nitrogen has an equivalence of I., III.

and v., we have by combination with oxygen several oxides of nitrogen. Thus:

N_2O or hyponitrous oxide.

N_2O_3, or nitrous oxide.



N_2O_5, or nitric oxide.

From these oxides, acids are formed on the plan of the water type by substituting for an atom of oxygen in each of the oxides two hydroxyls (OH); or by adding a molecule of water to each, Thus:—

N—O– H

N—O—H, $(HNO)_2 = N_2O + H_2O$, $(HNO)_2$ or hyponitrous acid—

$N_2O_3 + H_2O = (HNO_2)_2$, or nitrous acid,

$N_2O_5 + H_2O = (HNO_3)_2$ or nitric acid,

The salts of these acids are formed theoretically by replacing the hydrogen by basic elements.

HYPONITROUS OXIDE.

Formula, N_2O.

Preparation. Hyponitrous oxide is prepared by heating ammonium nitrate, (NH_4NO_3) which decomposes into hyponitrous oxide and water. Thus:—

Ammonium nitrate. Hyponitrous oxide. Water.

$$NH_4NO_3 = N_2O + H_2O.$$

Properties.

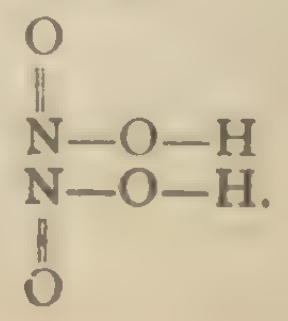
c. *Physical.* It is a colorless, odorless gas with a sweetish taste. It is commonly called "laughing-gas" or nitrous oxide. When inhaled it produces insensibility. It is heavier than air and soluble in water.

b. *Chemical.* It supports and accelerates combustion by being decomposed by heat and its oxygen set free.

Uses. As an anesthetic.

NITROUS ACID.

Formula. HNO_2



Properties. Nitrous acid is a very unstable compound. It is a bluish liquid, and has to be kept at a low temperature or it will decompose into nitric acid (HNO_3),water and nitrogen dioxide N_2O_2. Thus :—

Nitrous acid. Nitric acid Water. Nitric dioxide.

$$3HNO_2 = HNO_3 + H_2O + N_2O_2.$$

The salts formed from the union of nitrous acid and a base are called nitrites, and are much more stable compounds than the acid.

NITRIC ACID.

Formula, HNO_3, $\begin{matrix} O \\ \| \\ N—O—H. \\ \| \\ O \end{matrix}$

Preparation. Nitric acid is prepared by heating a mixture of sulphuric acid, (H_2SO_4,) and a nitrate, usually sodium nitrate ($NaNO_3$). The sulphuric acid unites with the sodium. Thus :—

Sod. nitrate. Sulph. acid. Hyd. sod. sulph. Nitric acid.

$$NaNO_3 + H_2SO_4 = HNaSO_4 + HNO_3.$$

Properties.

a. *Physical.* Nitric acid is a colorless, fuming, corrosive and poisonous fluid, heavier than water.

b. *Chemically* it is a powerful oxidizing agent. Many such substances as glycerine cotton, sugar, &c., are converted into explosives by the action of nitric acid. Mixed with hydrochloric acid (HCl), in the proportion of 1 to 3 it will dissolve gold, hence this mixture is called *aqua regia*.

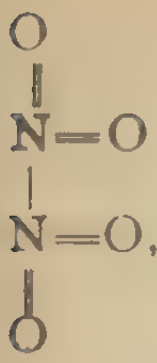
When nitric acid contains some of the nitrous acid it is then called *fuming* nitric acid, for the nitrous acid decomposes and sets free nitrogen dioxide, which combining with the oxygen of the air forms yellowish fumes of nitrogen teroxide (N_2O_4)

The union of nitric acid with bases forms nitrates, such as the nitrate of potash (KNO_3.)

Uses. Nitric acid is used for etching, for the manufacture of explosives, and for oxidizing purposes. Besides the regular oxides of nitrogen already considered, there are two others in which nitrogen having the equivalence of 3 and 5—two of the bonds saturate each other. Thus :—

$$\begin{matrix} N{=}O \\ | \\ N{=}O \end{matrix}$$

N_2O_2 or nitrogen dioxide.



N_2O_4 or nitrogen teroxide.

NITROGEN DIOXIDE.

Formula, N_2O_2.

Preparation. It is prepared by reducing nitric acid HNO_3, by metals as copper, silver, &c. Thus:

Copper. Nitric acid. Copper nitrate. Water. Nitrogen diox.

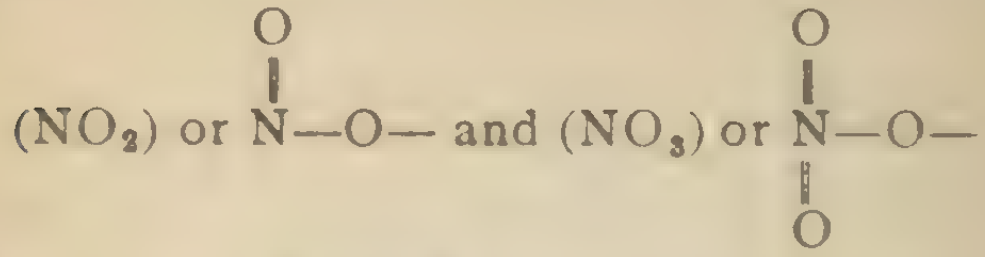
$$Cu_3 + (HNO_3)_8 = 3Cu(NO_3)_2 + 4H_2O + N_2O_2$$

Properties. It is a colorless gas and heavier than air. It does not support combustion, but has strong attractions for oxygen.

It will unite with the oxygen of the air, forming the yellowish brown suffocating fumes of *nitrogen ter-oxide* (N_2O_4). This is the vapor which arises from the chemical batteries using nitric acid and has such a disagreeable odor.

As nitrogen has an equivalence of I, III

and v, and oxygen of II, then (NO) or N—O—,



may all be considered as compound radicals, with one free bond and by union with basic elements form salts. Thus:

KNO or (NO)—K is potassium hyponitrite.

$NaNO_2$, or (NO_2)—Na is sodium nitrite, and

$$Ca(NO_3)_2 \text{ or } \begin{matrix}(NO_3)\\(NO_3)\end{matrix} > Ca$$ (calcium having two bonds), is calcium nitrate.

CHAPTER IV.

CARBON.

Symbol C. *Atomic weight* 12. *Equivalence* II. and IV.

History. Carbon in one form or another has been known for centuries.

Occurrence. Carbon in the free state exists only as a solid, but in three distinct forms.

1. Diamond,
2. Graphite,
3. Coal.

Chemically, these substances are the same but physically entirely different.

In the combined state carbon exists in the air, limestone, petroleum and in all animal and vegetable tissues.

DIAMOND.

History and *occurrence.* The diamond was found to consist only of carbon by Lavoisier, in 1776, who burned it in oxygen and

obtained as a product of this combustion carbon dioxide (CO_2).

It is found principally in India and Brazil.

Preparation. Diamonds have never been prepared physically or chemically.

Properties.

a. *Physical.* The diamond is a crystalline, transparent and solid body with a very high refractive and dispersive power. It is the hardest substance known, and a very poor conductor of heat and electricity.

b. *Chemically* little is known of the diamond except that it is combustible at a high temperature, and composed of carbon alone.

Uses. Diamonds are used as ornaments and for cutting hard substances, such as glass.

GRAPHITE.

Occurrence. Graphite is found in the United States, England, Ceylon and many other places as pure carbon.

Properties.

a. *Physical.* It is a crystallizable, soft, blackish solid, friable and oily to the touch. It is a good conductor of heat and electricity, and has a metallic lustre.

b. *Chemically* it is combustible only at high temperatures and with some difficulty, and is composed entirely of carbon.

Uses. It is used as a polish for metals, such as iron, preventing the formation of rust. It is also used in the manufacture of lead pencils and crucibles.

COAL.

Carbon as coal exists in two varieties.

a. Charcoal.

b. Mineral coal.

CHARCOAL.

Occurrence. Charcoal is not found free in nature, but is the result of destructive distillation.

Preparation, It is obtained by heating animal or vegetable matter to redness in nearly closed vessels. In this way the volatile matters are driven off, leaving the carbon behind.

Properties.

a. *Physical.* Charcoal is a black porous, soft solid and very light. It will absorb gases to a great extent.

b. *Chemically* it is a good oxidizer and very combustible.

Uses. It is used largely as a disinfectant by absorbing and oxidizing noxious gases, also as a decolorizing agent in refining sugar, and as a fuel.

MINERAL COAL.

Occurrence. It is found in nature as anthracite, bituminous and cannel coal, the former being much the purest carbon. It is the result of geological changes taking place in primitive vegetation, and is thus the product of vegetable life—the anthracite variety is of the oldest formation.

Properties.

a. *Physical.* Coal is an amorphius, hard, black, heavy solid with more or less luster.

b. *Chemically*, it is combustible with some difficulty, and is not composed of pure carbon.

Uses. Coal is used principally for heating and illuminating purposes.

All three forms of carbon are alike in being non-volatile, solid and remaining unaltered in the air at ordinary temperatures. They are not subject to decay,

COMPOUNDS.

I. OF CARBON AND HYDROGEN.

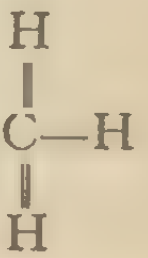
These compounds are called hydrocar-

bons, and many of them belong exclusively to organic chemistry.

HYDROGEN CARBIDE.

Synonyms. Methane, marsh gas and fire-damp.

Formula, CH_4 or



Occurrence. It occurs free in nature as "marsh gas" by the decomposition of vegetable matter under water, and also in coal mines as fire damp.

Preparation. It is chemically prepared by heating sodium acetate, ($NaC_2H_3O_2$) with caustic soda (NaOH). Thus :—

Sod. acetate. Caustic soda. Sod. carbonate. Methane.

$$NaC_2H_3O_2 + NaOH = Na_2CO_3 + CH_4.$$

Properties.

a. *Physical.* It is a colorless, odorless and tasteless gas. Slightly soluble in water and next to hydrogen in lightness.

b. *Chemically,* it is combustible, burning with a slightly luminous flame. With air it forms an explosive mixture. Methane

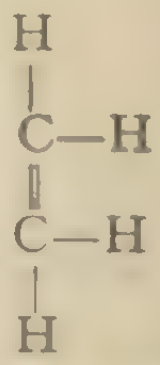
forms the first of a homogeneous series of organic compounds with the general formula of $C_nH_n + 2$—that is, having twice as many hydrogen atoms plus 2 than carbon.

Uses. It is used in some localities as an illuminating agent.

HYDROGEN DICARBIDE.

Synonyms. Carburetted hydrogen, olifeant gas and ethylene.

Formula. C_2H_4 or



Occurrence. It has been found free in nature among the gases in coal mines.

Preparation. It is prepared by the destructive distillation of coal in making illuminating gas.

Chemically, it is prepared by heating alcohol, C_2H_6O, with sulphuric acid, which splits up the alcohol into water and ethylene. Thus:

Alcohol. Sulphuric acid. Ethylene. Water. Sulph. acid.

$$C_2H_6O + H_2SO_4 = C_2H_4 + H_2O + H_2SO_4.$$

Properties.

a. *Physical.* It is a colorless gas with a sweetish taste and odor. It is soluble in water and lighter than air.

b. *Chemical.* It is combustible, burning with a brilliant white, but smoky, flame. When mixed with oxygen it is very explosive.

Uses. For illuminating purposes principally.

DIHYDROGEN DICARBIDE.

Synonym, Acetylene.

Formula, H_2C_2.

Preparation. Acetylene is prepared by the direct union of its constituents at a high temperature, as when the carbon points of the electrodes of a powerful chemical battery are brought near together in an atmosphere of hydrogen. It is produced in all cases of incomplete combustion.

Properties.

a. *Physical.* It is a colorless gas, soluble in water, and has a very disagreeable odor.

b. *Chemically*, it is combustible, burning with a bright, smoky flame.

It combines readily with some metals, as

copper and silver, thus forming violently explosive compounds.

2. OF CARBON AND OXYGEN.

CARBON DIOXIDE.

Formula. CO_2 or $C{=}O$ with O below.

History. In 1520 Paracelus discovered this gas, and it was the first one distinguished from the air.

Occurrence. It is found free in the air, especially in volcanic regions, and as it is a regular product of combustion and fermentation it is found in large quantities where such processes are going on.

It is often found in large quantities in low places, as at the bottom of wells, and called *choke damp.* In combination it exists in large quantities in lime stone.

Preparation. It is always the product of combustion in presence of oxygen. Thus :—

Carbon. Oxygen. Carbon dioxide.

$$C + O_2 = CO_2.$$

Chemically it is prepared by the action of

acids, as sulphuric (H_2SO_4) on calcic carbonate ($CaCO_3$). Thus :—

Calc. carbonate. Sulph. acid. Calc. sulphate. Water.

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2O$$

Carb. dioxide.

$$+ CO_2.$$

It is produced in large quantities by burning calcic carbonate ($CaCO_3$). Thus :—

Calc. carbonate. Calcic oxide. Carbon dioxide.

$$CaCO_3 = CaO + CO_2.$$

Properties.

a. *Physical.* Carbon dioxide is a colorless, odorless gas with an acid taste, very soluble in water, and 1½ times heavier than air. It is fatal to life if breathed in large quantities.

b. *Chemically* it will not support combustion, and is non-combustible. By its solution in water it unites with a molecule of water to form carbonic acid (H_2CO_3). Thus :—

Carbonic dioxide. Water. Carbonic acid.

$$CO_2 + H_2O = H_2CO_3.$$

This acid is readily decomposed by heat into H_2O and CO_2.

The salts of carbonic acid are carbonates, and are formed by replacing the hydrogen

atoms by basic elements as calcium, potassium, sodium, &c.

Uses. Soda waters and effervescent drinks are waters charged with carbon-dioxide.

CARBON MONOXIDE.

Formula, CO.

Preparation. It is produced by incomplete combustion, or by passing carbon-dioxide over heated carbon. Thus :—

Carbon dioxide. Carbon. Carbon monoxide.

$$CO_2 + C = 2CO.$$

Properties.

a. *Physical.* It is a colorless gas with a suffocating odor, slightly soluble in water and very fatal to animal life.

b. *Chemical.* It is combustible, burning with a blue flame, but will not support combustion. The gas will pass through heated cast-iron and in this way find egress into rooms from stoves or furnaces.

3. OF CARBON AND NITROGEN.

CYANOGEN.

Formula, $(CN)_2$ or $\begin{matrix} C{-}N \\ \| \\ C{-}N. \end{matrix}$

Symbol. Cy.

History. It was the first compound radical isolated.

Preparation. It is prepared by heating cyanide of mercury ($HgCy_2$,) Thus :—

Mercuric-cyanide. Mercury. Cyanogen.

$$HgCy_2 = Hg + Cy_2.$$

Properties.

a. *Physical.* It is a colorless, poisonous gas with an odor resembling that of peach blossoms.

b. *Chemical.* It is combustible, burning with a bluish flame. Although we have cyanogen existing as such by saturation of its own bonds still cyanogen (CN) is really a compound radical with an equivalence of one. Thus by its union with potassium there results potassium cyanide (K—(CN), with hydrogen, hydrogen cyanide (HCN), commonly called hydrocyanic acid or prussic acid.

HYDROCYANIC ACID.

Formula. HCN or H—(CN).

Preparation. Prussic acid can be prepared pure by passing sulphurated hydro-

gen (H_2S) over dry mercuric cyanide, Hg-$(CN)_2$. Thus :—

Mer'y cyanide. Sulph. hydrogen. Pruss. acid. Mer'y sulphide.

$$Hg(CN)_2 + H_2S = 2HCN + HgS.$$

Properties.

Prussic acid is a volatile liquid with an odor like that of bitter almonds, which contains a little of the acid. It is the most deadly poison known.

There are a great many cyanogens, but they belong generally to the organic compounds, and hence treated of under Organic chemistry.

CHAPTER V.

COMBUSTION.

Combustion in the strictest sense is oxidation, but commonly speaking it is rapid oxidation or oxidation accompanied by heat and light.

Products of combustion.

The products of combustion are of two kinds,

Physical.

Chemical.

a. The *physical* products are heat and light.

b. The *chemical* products are always water (H_2O) and carbon di-oxide (CO_2), the result of hydrogen and carbon burning in presence of oxygen.

ILLUMINATING GAS.

History. In 1792 Murdoch made gas illumination a success, it having been observed before this that the gases obtained from the distillation of coal were combustible.

Preparation. Illuminating gas is ordinarily prepared by heating bituminous coal to a high temperature, thus distilling off the combustible gases with their impurities.

The gas is first passed through pipes, thus condensing the water, gas-tar and ammonia liquors. It is next passed through the "purifier," which is filled with dry slaked lime (CaO), separating the other impurities from the gas, as sulphur compounds and carbon di-oxide (CO_2).

The gas, now freed from impurities, is ready for use. Gas can also be prepared from wood, petroleum, peat, &c. That prepared from coal is sometimes passed over a carbonator, an apparatus by means of which heavy hydrocarbons are added to the gas from naptha to make it more luminous.

Composition of illuminating gas. The principal gases constituting illuminating gas are:—

Marsh gas, CH_4.
Olifeant gas, C_2H_4.
Hydro carbons, C_nH_{2n}.

Carbon monoxide, CO.

Hydrogen, H.

There are always some impurities, even after the gas has been purified, such as carbon-dioxide, oxygen and sulphur compounds to a small per cent.

Properties.

a. *Physical.* Illuminating gas is colorless, with a very disagreeable suffocating odor. It is lighter than air and poisonous to animal life.

b. *Chemically* it is combustible, burning with a bright yellow flame. The collateral products from coal gas manufacture are:—

Coke,

Gas-tar,

Ammoniacal liquors.

Uses. The gases are used for illuminating purposes, the coke for fuel, the ammoniacal liquors for the manufacture of ammonium salts, and the gas tar for making anilines, phenols and hydro carbons.

ANIMAL AND VEGETABLE ECONOMY.

Animals as a rule produce in general those chemical substances which plants live

upon, and plants produce in their turn articles of food for animals. Thus :—

1. Animals produce and plants eat—

Carbon-dioxide, CO_2.

Water, H_2O.

Ammonia, NH_3

Salts, especially salts of, K, Na & Ca.

2. Plants produce and animals eat—

Vegetable products,

Oxygen.

CHAPTER VI.

HALOGEN GROUP.

Chlorine, bromine, iodine and fluorine form a group of elements which resemble each other to a marked degree, both physically and chemically.

They all form salts that contain no oxygen and resemble sea salt, hence called *halogen salts* and the elements *haloids*.

The ratio between their atomic weights, especially of the chlorine, bromine and iodine, is nearly constant.

CHLORINE.

Symbol, Cl. *Atomic weight*, 35.5. *Equivalence*, I.

History. Chlorine was discovered by Sheele in 1774.

Occurrence. Chlorine never occurs free in nature but always in combination. It is usually combined with potassium, sodium, magnesium and other metallic elements. It exists largely in mineral springs and sea

water. In the form of common salt (NaCl) it forms vast deposits in the earth's crust.

Preparation. It is largely prepared by the action of sulphuric acid (H_2SO_4) upon a mixture of common salt (NaCl) and manganese dioxide (MnO_2). Thus :—

Mang. diox. Sod. chloride. Sulphuric acid. Sod. sulphate

$$MnO_2 + (NaCl)_2 + 2H_2SO_4 = Na_2SO_4.$$

Manganic sulphate. Chlorine. Water.

$$+ MnSO_4 + Cl_2 + (H_2O)_2.$$

or by the action of hydrochloric acid (HCl) upon manganic oxide (MnO_2). Thus :—

Manganic oxide. Hydrochloric acid. Manganous chloride.

$$MnO_2 + 4HCl = MnCl_2$$

Water. Chlorine.

$$+ (H_2O)_2 + Cl_2.$$

Properties.

a. *Physical.* Chlorine is a yellowish green gas with a suffocating odor and astringent taste. It is very irritating to the air passages when inhaled, very soluble in water and much heavier than air.

b. *Chemically* it is very active, combining directly with all elements except nitrogen, oxygen and carbon.

It combines with free hydrogen in the presence of sunlight to form hydrochloric acid gas (HCl), and with powdered copper,

arsenic, antimony and phosphorous with the evolution of heat and light. Chlorine is combustible in hydrogen gas, but not in oxygen. Chlorine is sometimes called an *allotropic* element—that is, being capable of existing in two different physical conditions—for in the dark it is much less active than in the light. Free chlorine is a good bleaching agent for owing to the attraction of chlorine for hydrogen, water is decomposed, and the oxygen set free which destroys vegetable coloring matters by oxidizing them.

Uses. Chlorine is used as a bleaching agent and for a disinfectant. The so called chloride of lime is merely slacked lime impregnated with chlorine gas.

COMPOUNDS.

a. CHLORINE AND HYDROGEN.

HYDROGEN CHLORIDE OR HYDROCHLORIC ACID.

Formula. HCl or H—Cl.

Preparation. Hydrogen chloride may be prepared by direct union of its constituents by mixing equal volumes of chlorine and hydrogen gases, and exposing them to the sun-light. It is commercially made by the

action of sulphuric acid (H_2SO_4) on common salt (NaCl). Thus :—

Sod. chloride. Sulphuric acid. Sod. sulphate.

$$(NaCl)_2 + H_2SO_4 = Na_2SO_4$$

Hydrochloric acid.

$$+ (HCl)_2$$

Properties.

a. *Physical.* Hydrogen chloride is a colorless, pungent acid gas turning blue litmus red, very soluble in water forming hydrochloric acid.

b. *Chemically*, the gas is non-combustible.

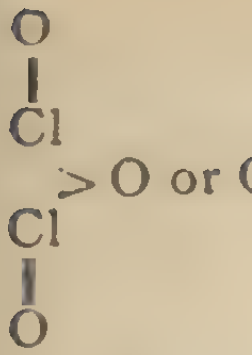
The liquid acid dissolves many of the metallic oxides as antimony, iron, tin, zinc, &c., forming chlorides.

Uses. Hydrochloric acid is used principally for chemical purposes.

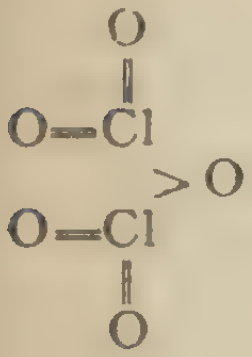
b. OF CHLORINE AND OXYGEN.

As chlorine may be a monad, triad, pentad or heptad, its union with the monad oxygen will result in the formation of four oxides. Thus :—

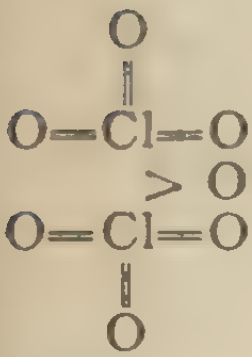
Cl
> O or Cl_2O, hypochlorous oxide.
Cl



> O or Cl_2O_3, chlorous oxide.



> O or Cl_2O_5, chloric oxide.



> O or Cl_2O_7, perchloric oxide.

From these oxides, acids are formed on the plan of the water type by substituting two hydroxyls (OH) for an atom of oxygen in each of the oxides, or which results in the same formula, adding one molecule of water. Thus :—

Cl—OH.

Cl—OH or $(HClO)_2 = Cl_2O + H_2O$ (H ClO)$_2$ hypochlorous acid.

$Cl_2O_3 + H_2O = (HClO_2)_2$ or chlorous acid.

$Cl_2O_5 + H_2O = (HClO_3)_2$ or chloric acid.

$Cl_2O_7 + H_2O = (HClO_4)_2$ or perchloric acid.

The salts of the acids are formed theoretically as usual by replacing the hydrogens with basic elements.

HYPOCHLOROUS OXIDE AND ACID.

Formula. Cl_2O.

Preparation. It is prepared by passing free chlorine over mercuric oxide HgO.

Mercuric oxide. Chlorine. Mercuric chloride

$$HgO + Cl \qquad HgCl_2$$

Hypochlorous oxide.

$$+ Cl\ O.$$

Properties.

a. *Physical.* It is a yellowish gas with an odor like free chlorine—very soluble in water.

b. *Chemically*, it is a very unstable compound, decomposing easily into chlorine and oxygen. It combines with water to form *hypochlorous acid*, (HClO) which is a very powerful bleaching agent.

The salts of this acid are called hypochlo-

rites, and are used for bleaching agents as for instance the calcium hypochlorite, $Ca(ClO)_2$ or

$$Ca < \begin{matrix} ClO \\ ClO. \end{matrix}$$

CHLOROUS OXIDE AND ACID.

Formula. Cl_2O_3.

Preparation. Chlorous oxide is prepared by reducing the chlorates by action of nitric acid (HNO_3) in presence of a reducing agent as arsenious acid.

Properties.

a. *Physical.* It is a greenish gas with a suffocating odor, and very soluble in water.

b. *Chemically*, it is a good bleaching agent and unites with water, forming *chlorous acid*, ($HClO_2$).

This compound is very unstable, as are its corresponding salts the chlorites, as potassium chlorite ($KClO_2$).

CHLORIC ACID.

Formula. $HClO_3$.

Preparation. Chloric acid is prepared by

the action of sulphuric acid (H_2SO_4) upon barium chlorate $Ba''(ClO_3)_2$. Thus :—

Barium chlorate. Sulphuric acid. Barium sulphate.

$$Ba(ClO_3)_2 + H_2SO_4 = BaSO_4$$

Chloric acid.

$$+ 2HClO_3.$$

Properties. Chloric acid is a very powerful oxidizing agent. It will set paper on fire when dropped upon it. The salts of this acid form chlorates and are also strong oxidizing agents.

Uses. The acid and its salts are used as oxidizing agents and in making pyrotechnical mixtures. Potassium chlorate is used in making oxygen.

PERCHLORIC ACID.

Formula, $HClO_4$.

Preparation. Perchloric acid is prepared by heating perchlorate of potassium ($KClO_4$) with sulphuric acid (H_2SO_4).

Properties.

a. *Physical.* It is an oily, colorless, fuming liquid with an odor of chlorine.

b. *Chemical.* It is a powerful oxidizing agent, setting paper and wood on fire when dropped upon them. Chloric acid and its

corresponding salts are the most stable of the chlorine acids and salts.

BROMINE.

Symbol, Br. *Atomic weight*, 80. *Equivalence*, I.

History. Bromine was discovered in the water of the Mediterranean Sea in 1826 by Ballad.

Occurrence. Bromine does not occur free in nature, but in combination with metallic elements, as sodium and magnesia, it is found in sea water and many mineral springs.

Preparation. It is prepared by evaporating water rich in bromides until a very concentrated solution remains, called mother liquid or bittern. By heating this bittern with manganese dioxide (MnO_2) and sulphuric acid (H_2SO_4) chlorine is set free from the decomposition of the chlorides present, which in turn sets free the bromine from the bromides.

Properties.

a. *Physical.* Bromine is a dark heavy liquid, with a suffocating odor.

b. *Chemically*, it is analogous to chlorine, but less active.

Uses. Bromine is used in photography and as a bleaching agent.

COMPOUNDS.

a. HYDROGEN AND BROMINE.

HYDROGEN BROMIDE OR HYDROBROMIC ACID.

Formula, HBr or H—Br.

Preparation. Hydrobromic acid is prepared by bringing phosphorus and bromine together in presence of water, when they unite with a loud explosion. Thus:

Phosphorus. Bromine. Water. Hydrobro. acid. Phosph. acid.

$$P + Br_5 + 4H_2O = 5HBr + H_3PO_4.$$

Properties.

a. *Physical.* Hydrobromic acid is a fuming, colorless gas with an acid reaction, and very soluble in water.

b. BROMINE AND OXYGEN.

Theoretically there are oxides and acids of bromine as there are of chlorine, but chemically there are not so many.

Thus there is hypobromous acid (HBrO) and bromic acid ($HBrO_3$), and from them their corresponding salts, formed by replacing the hydrogens by basic elements, as

calcium, forming calcium hypobromite ($Ca(BrO)_2$) and calcium bromate $Ca(BrO_3)_2$.

IODINE.

Formula, I. *Atomic weight*, 127. *Equivalence*, I.

History. Iodine was discovered in 1811 by Courtois, in sea water.

Occurrence. Iodine does not occur free in nature, but is combined with metals in sea water, and with sodium and magnesium in sea weed.

Preparation. Iodine is prepared by burning sea weed and dissolving this ash or kelp in water and thus obtaining the bittern or mother liquid. From this it is obtained by a process analogous to that of preparing bromine.

Properties.

a. *Physical.* Iodine is a black crystalline solid with a metallic lustre. Heated, it evolves the heaviest vapor known. Iodine is but slightly soluble in water, but readily in alcohol.

b. *Chemically* it is similar to bromine and chlorine, but less active than either. Free

iodine turns starch solutions a deep blue color.

Uses. Iodine is used principally in photography and medicine.

COMPOUNDS.

a. HYDROGEN AND IODINE.

HYDROIODIC ACID.

Formula, HI or H—I.

Preparation. This acid is prepared by acting upon phosphoric iodide (PI_3) with water. Thus :—

Phosphoric iodide. Water. Hydroiodic acid. Phosphorous acid

$$PI_3 + 3H_2O = 3HI + H_3PO_3.$$

Properties.

Hydroiodic acid is a colorless fuming acid gas, which is soluble in water.

b. OF IODINE AND OXYGEN.

As there are oxides and acids of chlorine and bromine so there are of iodine, the most important of which is *iodic acid* (HIO_3).

Preparation. Iodic acid may be prepared by oxidizing iodine with nitric acid (HNO_3) or by acting upon iodine water with chlorine gas. Thus :—

Iodine. Water. Chlorine. Iodic acid. Hydrochloric acid.

$$I + (H_2O)_3 + 5Cl = HIO_3 + 5HCl.$$

Properties.

Iodic acid is a colorless, crystalline solid, slightly soluble in water.

FLUORINE.

Symbol F. *Atomic weight* 19. *Equivalence* I.

Occurrence. Fluorine is found in nature combined with calcium, forming calcium fluoride (CaF_2) or "fluorspar;" with sodium (NaF); and aluminum (AlF_3), forming cryolite.

Preparation.

Fluorine has never been obtained in a free state.

Properties. Fluorine has never been known to combine with oxygen.

COMPOUNDS.

HYDROGEN AND FLUORINE.

HYDROFLUORIC ACID.

Formula, HF, or H-F.

Preparation. Hydrofluoric acid is prepared by acting upon fluorspar by sulphuric acid. Thus:

Fluoride of calcium, Sulph. acid. Calc. sulph. Hydrofluoric acid.

$$CaF_2 + H_2SO_4 = CaSO_4 + 2HF.$$

Properties.

a. *Physical.* It is a colorless, fuming acid gas, very irritating to the air tubes

when inhaled, and it is more or less corrosive.

b. *Chemically*, it has the power of attacking glass by uniting with the silicon to form silicon fluoride.

Uses. Hydrofluoric acid is used extensively in etching glass.

CHAPTER VII.

SULPHUR.

Symbol, S. *Atomic weight*, 32. *Equivalence*, II-IV.

History. Sulphur has been known from remote ages.

Occurrence. It occurs free as native sulphur in many volcanic regions.

In combination it occurs with iron, zinc, arsenic and many other mineral and metallic elements. It is also found in all animal and many vegetable tissues.

Preparation. It is prepared by roasting native sulphur or combined sulphur in furnaces, thus driving off the pure sulphur as a vapor which collects on the sides of the receiver as *flowers of sulphur*. Some of it is condensed into a liquid which is collected in moulds, where it hardens, forming *roll sulphur*.

Properties.

a. *Physical.* Sulphur occurs in three allotropic forms, viz.

1. The variety found in nature, which is a brittle crystalline solid, insoluble in water but soluble in carbon di-sulphide (CS_2).

2. The second variety is that obtained by the crystallization of sulphur fused at a high temperature.

It occurs as yellowish brown transparent crystals, insoluble in carbon di-sulphide, but changes slowly into variety (1).

3. The third variety is obtained by pouring melted sulphur into water; this forms a dark tenacious mass which can be drawn out into threads. It is insoluble in carbon di-sulphide, but slowly changes to variety (1).

b. *Chemical.* Sulphur is combustible at a temperature of 260°C, burning with a blue flame and forming sulphur dioxide (SO_2). Its vapor supports combustion, uniting with metals to form sulphides.

Uses. Sulphur is used for making sulphuric acid (H_2SO_4), and its vapor for bleaching straws.

COMPOUNDS.

I. SULPHUR AND HYDROGEN.

HYDROGEN SULPHIDE.

Formula. H_2S. $S<^{H}_{H}$

Occurrence. Hydrogen sulphide exists in volcanic gases and sulphur springs and is a product of decomposition.

Preparation. It can be prepared by direct union of its constituents by means of heat, but it is usually made by the action of sulphuric acid (H_2SO_4) on sulphide of iron (FeS). Thus:—

Sulphide iron. Sulph. acid. Sulphate iron. Sulphide hydrogen.

$$FeS + H_2SO_4 = FeSO_4 + H_2S.$$

Properties.

a. *Physical.* It is a colorless gas with an odor of rotten eggs. It is heavier than air and very soluble in water. It is quite poisonous to animal life.

b. *Chemically*, it is combustible, burning with a blue flame. It is easily decomposed by oxidizing agents. It unites with metallic bases and their oxides to form sulphides.

Uses. It is used in laboratories as a reagent, uniting with metallic elements in so-

lution to form insoluble sulphides, which have various colors.

2. OF CARBON AND SULPHUR.

CARBON DI-SULPHIDE.

Formula. CS_2, $C{=}S$ (‖ S)

Preparation. Carbon di-sulphide is best prepared by passing sulphur vapor over red hot charcoal. Thus:—

Carbon. Sulphur. Carbon di-sulphide.

$$C + S_2 = CS_2.$$

Properties.

a. *Physical.* It is a liquid with an etherial odor when pure, but when impure has a very disagreeable odor. It is not soluble in water.

b. *Chemical.* When vaporized it is combustible, burning with a bright blue flame. It has remarkable solvent properties, dissolving phosphorus, iodine, sulphur, resins, fats, &c.

Uses. It is used principally as a solvent and for extracting oils and fats.

3. OF SULPHUR AND OXYGEN.

As sulphur may be a diad, tetrad or hexad

there result from its union with oxygen the following oxides.

S=O, or SO, hyposulphurous oxide.

O
‖
S=O, or SO_2, sulphurous oxide.

O
‖
S=O, or SO_3, sulphuric oxide.
‖
O

From these oxides there are formed, on the plan of the water type, the following acids.

OH
S<
OH, or H_2SO_2, hyposulphurous acid.

H_2SO_3, sulphurous acid.

H_2SO_4, sulphuric acid.

Hyposulphurous acid and oxide are unknown chemically.

SULPHUROUS OXIDE AND ACID.

Formula. SO_2, S=O.
‖
O

Occurrence. Sulphurous oxide is found free in volcanic gases.

Preparation. It is always formed by the

combustion of sulphur in the presence of oxygen. Thus:

Sulphur. Oxygen. Sulphurous oxide.

$$S + O_2 = SO_2,$$

or by the reduction of sulphuric acid by copper. Thus:—

Copper. Sulphuric acid. Copper sulphate. Sulphurous oxide.

$$Cu + 2H_2SO_4 = CuSO_4 + SO_2$$

Water.

$$+ 2H_2O.$$

Properties.

a. *Physical.* Sulphurous oxide, or commonly called sulphur di-oxide, is a pungent gas with a suffocating odor. Twice heavier than air and freely soluble in water.

b. *Chemical.* It is not combustible or a supporter of combustion. It unites with metallic oxides to form sulphites. Combining with a molecule of water it forms sulphurous acid, which is a liquid having an acid reaction. It is a fine bleaching agent, owing to its affinity for oxygen, thus uniting with the oxygen of the water or coloring agents, decomposing the compounds and setting free hydrogen. Best bleaching agent for woolens and silks.

Sulphurous acid (H_2SO_3) is a reducing

agent, owing to its affinity for oxygen, which combines with the sulphurous acid to form sulphuric acid. By replacing one or more of the hydrogens in sulphurous acid by basic elements, sulphites are formed. When only one of two hydrogens is thus replaced the product is called an acid salt.

Uses. Sulphurous oxide is a good disinfectant for sick rooms. The acid and salts are used for bleaching woolens and preserving cider, fruits, &c.

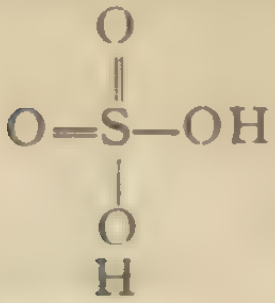
SULPHURIC OXIDE.

Formula, SO_3 or $O{=}S{=}O$ (with $\|$ O).

Preparation. Sulphuric oxide is prepared by oxidizing sulphurous oxide by means of heat.

Properties. It is a wax like solid—crystallizing, thus resembling asbestos. It melts at 16°C, forming the fluid oxide which has the highest co-efficient of expansion known. It unites with water with great heat, forming sulphuric acid (H_2SO_4).

SULPHURIC ACID.



Formula, H_2SO_4.

Occurrence. It occurs free in some mineral springs, and in the secretions of many animals. Combined with metallic elements it occurs in great abundance.

Preparation. It may be prepared by adding water to sulphuric oxide. Thus :—

Water. Sulphuric oxide. Sulphuric acid.

$$H_2O + SO_3 = H_2SO_4.$$

Commercially it is prepared by the four following steps.

1. Sulphur is burned in the air, thus producing sulphurous oxide. Thus :—

Sulphur. Oxygen. Sulphurous oxide.

$$S + O_2 = SO_2.$$

2. This sulphurous oxide (SO_2) is then mixed with nitrogen tetroxide, forming sulphuric oxide and nitrogen dioxide (N_2O_2.) Thus :—

Sulphuric oxide. Nitrogen tetroxide. Sulphuric oxide.

$$2SO_2 + N_2O_4 = 2SO_3$$

Nitrogen dioxide.

$$+ N_2O_2.$$

3. This sulphuric oxide (SO_3.) is then mixed with steam (H_2O), which unites with it, forming sulphuric acid (H_2SO_4). Thus :—

Sulphuric oxide. Water. Sulphuric acid.

$$SO_3 + H_2O = H_2SO_4.$$

4. The nitrogen tetroxide (N_2O_4) is obtained again for use by reoxidation of nitrogen dioxide (N_2O_2) from the air. Thus :—

Nitrogen dioxide. Oxygen. Nitrogen tetroxide.

$$N_2O_2 + O_2 = N_2O_4$$

Properties.

a. *Physical.* Sulphuric acid is a colorless, oily, corrosive liquid much heavier than water. It is decomposed at high temperatures, and has strong affinity for water.

It will extract moisture from the air, drying it; and from organic matter, charring it.

b. *Chemically*, one or more of its hydrogens may be replaced by basic elements forming sulphates, which are very stable compounds.

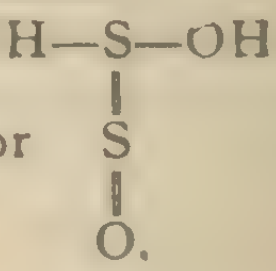
Uses. Sulphuric acid is principally used in the manufacture of chemicals, as nitric, hydrochloric, phosphoric and other acids, and commercial products.

THIONIC ACIDS.

There are a few sulphur acids—but of little importance—which are formed of two atoms of sulphur combining with oxygen and hydrogen instead of one. Thus :—

SULPHO-SULPHURIC ACID.

Formula, $H_2S_2O_3$ or



Properties.

The salts formed from this acid are called hyposulphites, the most important of which is sodium hyposulphite, $Na_2S_2O_3$—formed by replacing the hydrogen by the metallic element sodium.

Uses. This salt is used as a medicine and in photography. .

CHAPTER VIII.

SILICON AND BORON.

SILICON.

Symbol Si. *Atomic weight* 28. *Equivalence* IV.

History. Silicon was first obtained pure by Berzelius in 1823.

Occurrence. It does not occur free in nature but in combination with oxygen, aluminum, potassium, &c., it forms a large portion of the rocks making up the earth's crust.

Preparation. Silicon is prepared by the action of sodium (Na) upon potassium flo-silicate (K_2SiF_6). Thus: —

Pot. flo-silicate. Sodium. Pot. fluoride. Sod. fluoride. Silicon.

$$K_2SiF_6 + 4Na = 2KF + 4NaF + Si.$$

Properties. It is an amorphous, lustreless powder, very hard and melting only at a very high temperature. At a still higher temperature it takes fire forming by its combustion silicic oxide.

COMPOUNDS.

1. OF SILICON AND HYDROGEN.

HYDROGEN SILICIDE.

Formula, H_4Si, $H-\underset{\displaystyle H}{\overset{\displaystyle H}{\overset{|}{\underset{|}{Si}}}}-H$.

Preparation. It is prepared by the decomposition of magnesium silicide Mg_2S by hydrochloric acid, HCl. Thus :—

Mag. silicide. Hydrochloric acid. Mag. chloride.

$$Mg_2Si + 4HCl = 2MgCl_2$$

Hydrogen silicide.

$$+ H_4Si.$$

Properties. It is a colorless gas, burning with a white flame upon coming in contact with oxygen, thus forming white clouds of silica, (SiO_2) and water.

2. OF SILICON AND OXYGEN.

SILICA OR SILICON OXIDE.

Formula. SO_2.

Occurrence. Silica occurs abundantly in nature as crystalline quartz, opal, amethyst, agate, &c. It also occurs as sand-stone. In combination with metallic oxides forming silicates it is very abundant. It is found dissolved in many mineral waters, and as a

constituent of many animal and vegetable tissues.

Preparation. It is prepared by the oxidation of silicon or by the dehydration of silicic acid, H_2SiO_3.

Properties.

a. *Physical.* Silica occurs in two forms, the crystalline and amorphous. It is very hard and will scratch glass. It is but slightly soluble in water and only in one acid, hydrofluoric (HF). It is fusible but only at very high temperatures.

b. *Chemical.* Silica heated with sodium or potassium hydrate under pressure will dissolve and form "fluid glass," or sodium or potassium silicate. Thus :—

Potass. hydrate. Silica. Water. Potass. silicate.

$$2KOH + SiO_2 = H_2O + K_2SiO_3.$$

If an acid is added to this silicate it will decompose, and uniting with the hydrogen of the acid form *Silicic acid*, H_4SiO_4. This acid is a gelatinous substance, not very stable for it loses a molecule of water upon standing, forming what is called *meta*-silicic acid, H_2SiO_3.

By replacing one or more of the hydrogens of either acid by a basic element, sili-

cates are formed. The sodium and potassium meta-silicates are soluble in water.

Glass is a double silicate, and there are two principal varieties.

a. Flint glass, or potassio-lead silicate.

b. Crown or window glass, or sodio-calcic silicate.

BORON.

Symbol, B. Atomic weight, 11. *Equivalence,* III.

History. The salts of boron have been known for centuries, but pure boron was obtained in 1808.

Occurrence. It does not occur free in nature, but in combination with sodium, potassium and ammonia it is found in many mineral waters, and as mineral deposits in the earth's crust.

Preparation. It is prepared by the action of sodium (Na) on potassium flo-borate (KBF_4). Thus :—

Pot. flo-borate. Sodium. Pot. fluoride. Sod. fluoride, Boron.

$$KBF_4 + 3Na = KF + 3NaF + B.$$

Properties. Boron is a chocolate colored powder slightly soluble in water. By dissolving it in aluminum it will crystallize,

forming a solid having many of the physical properties of the diamond.

COMPOUNDS.

1. OF BORON AND OXYGEN.

BORIC OXIDE.

Formula, B_2O_3.

Preparation. It is produced by the combustion of boron or boracic acid, H_3BO_3 in presence of oxygen.

Properties. Boric oxide is a colorless, brittle, transparent solid.

2. OF BORON AND HYDROGEN.

BORIC OR BORACIC ACID.

Formula. H_3BO_3.

Occurrence. Boric acid occurs free in nature, especially in solution in the waters of volcanic regions.

Preparation. It is easily prepared by evaporating the waters holding it in solution.

Properties.

a. *Physical.* It occurs as white crystalline scales, soluble in water and having an acid reaction.

b. *Chemical.* By replacing one or more

of the hydrogens by basic elements, borate salts are formed.

Thus sodium borate (NaH_2BO_3) is formed by replacing one hydrogen. The hydrogen contained in this salt is combined with oxygen in the form of water, and can be driven off by heat. Borax or fused sodium borate ($Na_2B_4O_7$) is a combination of the dehydrated sodium borate ($NaH_2BO_3—H_2O$), and boric oxide B_2O_3. Thus:—

Dehydrated sod. borate, Boric oxide. Borax

$$2NaBO_2 + B_2O_3 = Na_2B_4O_7.$$

CHAPTER IX.

PHOSPHORUS.

Symbol, P. *Atomic weight*, 31. *Equivalence*, I-II-V.

History. Phosphorous was discovered by Brant in 1669, from urine.

Occurrence. Phosphorus does not occur free in nature, but usually in combination with calcium, magnesium and lead.

Preparation. Burnt bone is allowed to macerate in sulphuric acid (H_2SO_4) for some hours. This extracts from the bones phosphorus in the form of calcium meta-phosphate $Ca(PO_3)_2$, which remains in solution, and is poured off from the remainder of the mixture and evaporated to a syrupy consistency. It is next mixed with charcoal and heated to redness—the phosphorus distills over and is condensed. Thus :—

Calc. meta phosphate. Carbon. Calc. ortho-phosphate.

$$3Ca(PO_3)_2 + 10C = Ca_3(PO_4)_2$$

Carbonyl. Phosphorus.

$$+ 10CO + P_4.$$

Properties.

a. *Physical.* Phosphorus exists in two allotropic forms. In its usual form it is a colorless, transparent and waxy substance with a peculiar odor. It is insoluble in water but soluble in ether, alcohol, oils and carbon di-sulphide,

b. *Chemical.* It oxidizes readily and burns in the air, hence must be kept under water. Owing to slow combustion it is luminous in the dark and forms phosphorous oxide (P_2O_3). By rapid combustion it forms phosphoric oxide (P_2O_5). By heating this form of phosphorus to 250°C, it is converted into the other allotropic state.

This form of phosphorus is a chocolate-red powder insoluble in ordinary solvents, does not oxidize readily in the air, has no odor, is not poisonous, does not burn easily and will not shine in the dark.

Uses. Phosphorus is used principally in making matches—both forms of phosphorus being employed.

COMPOUNDS.

1. OF PHOSPHORUS AND HYDROGEN.

HYDROGEN PHOSPHIDE OR PHOSPHINE.

Formula, H_3P, $\begin{matrix} H \\ | \\ P-H \\ | \\ H \end{matrix}$

Preparation. Phosphine is usually prepared by boiling phosphorus with a solution of caustic potash (KOH). Thus :—

Phosphorus. Potass. hydrate. Water. Pot. hypophosphite.

$$4P + 3KOH + 3H_2O = 3KH_2PO_2$$

Phosphine.

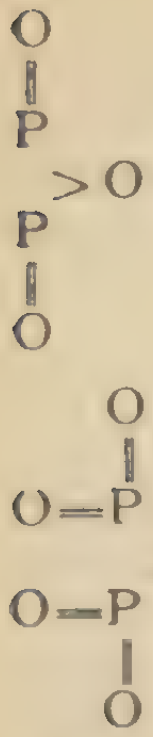
$$+ PH_3.$$

Properties. Phosphine is a colorless gas with a garlic like odor, slightly soluble in water, and burns with a brilliant flame.

2. OF PHOSPHORUS AND OXYGEN.

As phosphorus has an equivalence of I, III and V, by its union with oxygen, three oxides are formed. Thus :—

$\begin{matrix} P \\ > O \\ P \end{matrix}$ or P_2O, hypophosphorous oxide.



$> O$ or P_2O_3, phosphorous oxide.

$> O$ or P_2O_5, phosphoric oxide.

There are no known acids formed directly by the union of hydrogen with the hypo or ous oxide, but there are three well known acids formed from the ic oxide by substituting two hydroxyls for 1, 2, and 3 of the oxygen atoms, or in other words, by adding 1, 2 and 3 molecules of water. Thus :—

$P_2O_5 + H_2O = 2HPO_3$, called metaphosphoric acid.

$P_2O_5 + 2H_2O = H_4P_2O_7$, called pyrophosphoric acid.

$P_2O_5 + 3H_2O = 2H_3PO_4$, called orthophosphoric acid.

PHOSPHORIC OXIDE.

Formula. P_2O_5.

Preparation. Phosphoric oxide is prepared by the rapid combustion of phosphorus in the air. $P_2 + O_5 = P_2O_5$.

Properties. It is a white, amorphous powder, attracting moisture from the air and dissolving in water with a hissing noise, forming ortho-phosphoric acid (H_3PO_4).

ORTHO-PHOSPHORIC ACID.

Formula. H_3PO_4.

Preparation. Ortho-phosphoric acid is prepared by adding boiling water to phosphoric oxide (P_2O_5). Thus:

Phosphoric oxide. Water. Ortho-phosphoric acid.

$$P_2O_5 + 3H_2O = 2H_3PO_4,$$

or it can be prepared by the oxidation of phosphorus with nitric acid.

Properties. Phosphoric acid is a clear, syrupy liquid with an acid reaction.

As there are three hydrogens that can be replaced by basic elements it follows that there are three varieties of salts that can be formed, viz.—acid, normal and double (see theoretical chemistry), and that the salts may be mono, di or tri basic.

Thus calcium is a diad and will take the place of two hydrogens, hence $Ca''HPO_4$ is mono-calcic-ortho-phosphate and an acid salt. *Again*, di-calcic-ortho-phosphate must be an acid salt with a formula $Ca_2H_2(PO_4)_2$, for to saturate two atoms of calcium it will take four atoms of hydrogen, and as there are only three in the acid two molecules must be taken or $H_6(PO_4)_2$, to combine with the calcium. *Again*, tri-calcic-ortho-phosphate, or bone phosphate, is $Ca_3(PO_4)_2$ and a normal salt.

Ammonium magnesium-ortho-phosphate $(NH_4)'Mg''PO_4$ is a double salt of ortho-phosphoric acid.

PYRO-PHOSPHORIC ACID.

Formula. $H_4P_2O_7$.

Preparation. This acid is prepared by heating phosphoric acid, H_3PO_4, to 215°C and thus driving off a molecule of water. Thus :—

Ortho-phosphoric acid. Water. Pyro-phosphoric acid.

$$2H_3PO_4 - H_2O = H_4P_2O_7.$$

Properties. Pyro-phosphoric acid is a clear fluid substance with an acid reaction. By substituting for the hydrogens in this acid, basic elements, different kinds of pyro

salt are formed, as are ortho salts from ortho-phosphoric acid. Thus:—$Na_4P_2O_7$ is normal and $Na_2H_2P_2O_7$ acid sodium pyrophosphate.

META-PHOSPHORIC ACID.

Formula, HPO_3.

Preparation. Meta-phosphoric acid is prepared by heating ortho-phosphoric acid to redness, thus driving off a molecule of water.

Ortho-phosphoric acid. Water. Meta-phosphoric acid.

$$H_3PO_4 - H_2O = HPO_3,$$

or by dissolving phosphoric oxide (P_2O_5) in water. Thus:

Phosphoric oxide. Water. Meta-phosphoric acid.

$$P_2O_5 + H_2O = 2HPO_3.$$

Properties.

This acid is a transparent ice-like solid, soluble in water and with an acid reaction. As it contains but one hydrogen it forms but one class of salts—normal.

PHOSPHOROUS OXIDE.

Formula, P_2O_3.

Preparation. It is formed by the incomplete oxidation of phosphorus.

Properties. It is a volatile, white, flaky

solid with a peculiar odor. It dissolves in water with a hissing noise.

HYPOPHOSPHITES.

By adding three molecules of water to hypophosphorous oxide there is formed theoretically hypophosphorous acid H_3PO_2.

$$(P_2O + 3H_2O = 2H_3PO_2).$$

By substituting basic elements in place of the hydrogens, salts are formed called hypophosphites—used in medicine.

CHAPTER X.

ARSENIC.

Symbol, As. *Atomic weight*, 75. *Equivalence*, III–V.

History. Pure arsenic was first obtained by Schroeder, in 1694.

It had been known in combination for centuries.

Occurrence. Arsenic occurs free as native arsenic, but usually in combination with iron, cobalt, nickel and sulphur.

Preparation. It is prepared by roasting arsenic ore, thus driving off the volatile arsenic, which is condensed as arsenious oxide As_2O_3.

This oxide is then reduced by heating it with charcoal. Thus :—

Arsenious oxide, Carbon. Arsenic, Carbon monoxide

$$As_2O_3 + 3C = As_2 + 3CO.$$

Properties. Arsenic is a brittle solid with a metallic lustre. It is volatilized by heat

giving off yellowish fumes with a garlic like odor.

At a high temperature arsenic burns with a blue flame forming arsenious oxide, (As_2O_3).

COMPOUNDS.

1. OF ARSENIC AND HYDROGEN.

HYDROGEN ARSENIDE OR ARSINE.

Formula. H_3As, AsH_3.

Preparation. Arsenic is prepared by the action of sulphuric acid (H_2SO_4), upon zinc arsenide (As_2Zn_3). Thus:—

Zinc arsenide, Sulph. acid. Zinc sulphate, Arsine.

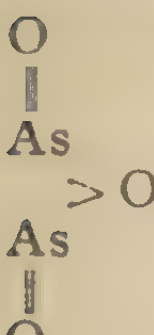
$$As_4Zn_3 + 3H_2SO_4 = 3ZnSO_4 + 2H_3As.$$

Properties. Arsine is a colorless gas with a garlic like odor. It is soluble in water and combustible, forming by its combustion arsenious oxide (As_2O_3).

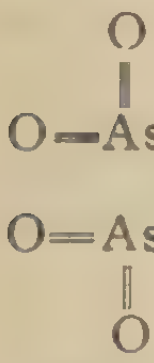
2. OF ARSENIC AND OXYGEN.

As arsenic has an equivalence of III or V,

two oxides are formed by its union with oxygen, Thus:—



$> O$ or As_2O_3, Arsenious oxide.



$> O$ or As_2O_5, Arsenic oxide.

From these oxides by the addition of 3 molecules of water, two acids are formed.

$As_2O_3 + 3H_2O = 2H_3AsO_3$, Arsenious acid, and—

$As_2O_5 + 3H_2O = 2H_3AsO_4$, Arsenic acid.

ARSENIOUS OXIDE.

Formula. As_2O_3.

Synonym, arsenious acid.

Occurrence. It occurs free in nature under the name of arsenolite.

Preparation. It is prepared by roasting

arsenic ores in an access of air, and condensing the vapor formed.

Properties. Arsenious oxide exists in two forms, crystalline and vitreous. The form taken by the oxide seems to depend upon the amount of heat the vapor has been subjected to in condensing. If less than 250°C, the octahedral or rhomboid crystals are formed, if more than that, the vitreous variety.

The crystalline form by exposure to the air at ordinary temperatures slowly changes to the vitreous variety.

Both forms are soluble in alkaline solutions, hydrochloric acid (HCl,) and slightly in water. From its solution it crystallizes in octahedral crystals.

ARSENIOUS ACID.

Formula. H_3AsO_3.

Preparation. It is obtained by dissolving arsenious acid in water.

Properties. Arsenious acid is a styptic liquid, with an acid reaction, not very stable.

Arsenite salts are formed by substituting for the hydrogen in the acid, basic elements. These salts are not easily decom-

posed. Among the most important of them are potassium arsenite (K_3AsO_3), which is the principal ingredient in *Fowler's solution* of arsenic, *Sheele's green* and other pigments.

Paris green is a double salt of arsenite and acetate of copper.

ARSENIC OXIDE.

Formula. As_2O_5.

Preparation. This is prepared by heating arsenious oxide (As_2O_3) with nitric acid (HNO_3), evaporating the solution to dryness, and heating to 270°C.

Properties. It is an opaque, white, amorphous solid, and is decomposed by heat. By its solution in water forming *Arsenic acid* (H_3AsO_4).

Arsenate salts are formed from this acid by substituting for the hydrogen basic elements.

All forms of arsenic are poisonous, especially the *ite* salts.

Tests for arsenic.

a. *Sublimation test :—*

Heat arsenic in a glass tube with air. Crystals of arsenious oxide, As_2O_3 will be deposited on the glass.

b. *Reduction test :—*

Heat arsenic in a closed tube with charcoal. A metallic mirror of arsenic will form on sides of tube.

c. *Sulphurretted hydrogen test :—*

Pass sulphurretted hydrogen (H_2S) through a hydrochloric acid (HCl) solution of arsenic. A lemon yellow precipitate will form, which is soluble in ammonia.

d. *Sulphate of copper test :—*

Add to an ammoniacal solution of arsenic, a solution of sulphate of copper ($CuSo_4$). A green precipitate is produced, which is soluble in access of the ammonia.

e. *Marsh's test :—*

Add a solution of arsenic to sulphuric acid (H_2SO_4), and then add metallic zinc. Collect the gas given off which will contain hydrogen arsenide. This gas will have the properties of hydrogen arsenide. If the gas is ignited and the flame directed against a piece of cold porcelain, a black stain will remain which is soluble in hyposulphite of soda (Na_2SO_2).

CHAPTER XI.

ANTIMONY AND BISMUTH.

ANTIMONY, (STIBIUM).

Symbol, Sb. *Atomic weight*, 122. *Equivalence*, III and V.

History. Pure antimony was obtained by Valentine, in 1492.

Occurrence. It occurs in nature, both free and combined with oxygen, sulphur, silver, &c.

Preparation. It is prepared by melting sulphide of antimony (Sb_2S_3) with iron. Thus:—

Sulphide of antimony, Iron. Sulph. of iron, Antimony.

$$Sb_2S_3 + 3Fe = 3FeS + 2Sb.$$

or by roasting antimony ore thus obtaining the oxide (Sb_2O_3) which is reduced by heating with charcoal, analogous to the preparation of arsenic.

Properties. Antimony is a bluish white, brittle metal not easily tarnished. It fuses and burns at elevated temperatures producing by its combustion antimonous oxide (Sb_2O_3).

Uses. Antimony is used largely in making type metal.

COMPOUNDS.

1. OF ANTIMONY AND HYDROGEN.

HYDROGEN ANTIMONIDE.

Formula. H_3Sb, SbH_3.

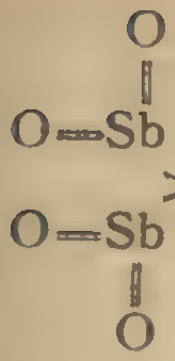
Preparation and *properties.* It is prepared in a manner analogous to that of hydrogen arsenide (AsH_3), and has properties very similar to that compound. Its stain on porcelain, however, is darker and not soluble in hyposulphite of soda (Na_2SO_3)

2. OF ANTIMONY AND OXYGEN.

As antimony has an equivalence of III and V, by its union with oxygen two oxides are formed. Thus :—

Sb_2O_3, antimonous oxide.

Sb_2O_5, antimonic oxide.



> O, Sb_2O_5, antimonic oxide.

And from these oxides, as in arsenic, *ous* and *ic* acids are formed. Chemically, the oxides, acids and salts of antimony resemble those of arsenic. The principal oxide of antimony is the *ous* oxide.

ANTIMONOUS OXIDE.

Formula, Sb_2O_3.

Occurrence. This oxide occurs free in nature.

Preparation. It may be prepared by boiling antimonous chloride ($SbCl_3$) with a solution of sodium carbonate (Na_2CO_3).

Antimonous chloride. Sodium carbonate. Antimonous oxide.

$$2SbCl_3 + 3Na_2CO_3 = Sb_2O_3$$

Sod. chloride. Carbon dioxide.

$$+ 6NaCl + 3CO_2.$$

Properties. It is a dirty white powder slightly soluble in water, forming the *ous* acid.

Salts corresponding to this acid are used in medicine. Tartar emetic ($2KSbO, C_4H_4$-

O_6) is a double salt of potassium antimony tartrate.

3. ANTIMONY AND SULPHUR.

There are two sulphides of antimony, the *ous* (Sb_2S_3) and *ic* (Sb_2S_5). The former is found free in nature.

BISMUTH.

Symbol, Bi. *Atomic weight*, 210. *Equivalence*, III and V.

History. It was discovered pure by Valentine in 1492.

Occurrence. Bismuth occurs free in nature, but usually in combination with oxygen or sulphur.

Preparation. It is prepared by reducing the sulphide of bismuth (Bi_2S_3) with charcoal at a high temperature.

Properties. Bismuth is a hard, brittle, reddish white metal, remaining untarnished in dry air. At high temperatures it melts and burns with a white flame, forming by its combustion bismuthous oxide (Bi_2O_3).

Uses. Bismuth is used principally in forming alloys.

COMPOUNDS.

1. OF BISMUTH AND OXYGEN.

As bismuth has an equivalence of III and V, by its union with oxygen two oxides are formed, the *ous* (Bi_2O_3) and *ic* (Bi_2O_5), and from them the corresponding acids and salts.

2. OF BISMUTH AND CHLORINE.

Chlorine being a monad unites with bismuth to form two compounds--bismuthous chloride ($Bi^{III}Cl_3$) and bismuthic chloride ($Bi^{V}Cl_5$). In the same way are formed two compounds of bismuth and sulphur—the *ous* (Bi_2S_3), which occurs free in nature, and the *ic* (Bi_2S_5).

CHAPTER XII.

TIN, ZINC, CADMIUM.

TIN (Stannum).

Symbol, Sn. *Atomic weight*, 118. *Equivalence*, II and IV.

Occurrence. Tin occurs principally as stannic oxide (SnO_2) called cassiterite.

Preparation. This stannic oxide is roasted and then melted with charcoal, which combines with the oxygen leaving the metallic tin.

Properties. Tin is a soft white metal, very malleable and ductile. It has a peculiar odor, crackles (tin cry) when broken, and is a good conductor of heat and electricity. It is not volatile, but burns at a high temperature, forming stannic oxide (SnO_2).

Uses. Tin is used in making foil and alloys.

COMPOUNDS.

I. OF TIN AND CHLORINE.

As tin has an equivalence of two or four

it forms by its union with chlorine two compounds. Thus:

$$Sn\begin{cases} Cl \\ Cl \end{cases}$$ or $SnCl_2$, Stannous chloride.

$$Cl-Sn(-Cl)-Cl$$ or $SnCl_4$, Stannic chloride.

STANNOUS CHLORIDE.

Formula, $SnCl_2$.

Preparation. It is prepared by the action of mercurous chloride, Hg_2Cl_2, on tin filing. Thus:—

Mercurous chloride. Tin. Mercury. Stannous chloride.

$$Hg_2Cl_2 + Sn = Hg_2 + SnCl_2.$$

Properties. This is a grayish white solid, soluble in water. Upon evaporating its aqueous solution, crystals, known as "tin salt," are formed, which are used in dyeing.

STANNIC CHLORIDE.

Formula, $SnCl_4$.

Preparation. It is prepared by distilling mercuric chloride, $HgCl_2$, with tin. Thus:—

Mercuric chloride. Tin. Mercury. Stannic chloride.

$$2HgCl_2 + Sn = Hg_2 + SnCl_4.$$

Properties. It is a colorless fuming liquid

uniting directly with alkaline chlorides to form compounds.

It is used in dyeing.

2. OF TIN AND OXYGEN.

As there are two compounds of tin with chlorine, so two are formed by the union of tin with oxygen.

STANNOUS OXIDE.

Formula, SnO.

Properties. This oxide is a black powder which absorbs oxygen from the air, forming stannic oxide (SnO_2).

STANNIC OXIDE.

Formula, SnO_2.

Preparation. This occurs free in nature but mixed with impurities. It is best prepared by burning tin in presence of oxygen.

Properties. Thus prepared it is a very hard white powder, insoluble in all acids save hydrofluoric (HF). It is used for polishing glass under the name of "putty powder."

By the addition of a molecule of water to each oxide two corresponding acids are formed called stannous (H_2SnO_2) and stannic (H_2SnO_3) acids. Stannite and stannate

salts are formed from the acids by substituting basic elements for the hydrogens of each.

3. OF TIN AND SULPHUR.

The principal sulphide of tin is the stannic (SnS_2). It is a golden yellow powder known as "mosaic gold."

ZINC.

Symbol, Zn. *Atomic weight*, 65. *Equivalence*, II.

Occurrence. It occurs in nature as zinc ore in combination with oxygen, silica, carbon, sulphur, &c.

Preparation. It is prepared by first roasting the ore and then distilling with charcoal. The oxygen is removed by combining with the carbon, and the zinc volatilizing, distills over and is condensed.

Properties. Zinc is a bluish white, brittle metal. It is both ductile and malleable, oxidizes readily. It volatilizes and burns at high temperatures, forming zinc oxide (ZnO).

Uses. Zinc is so readily acted upon by acids it forms the positive element in nearly all voltaic batteries.

It is also largely used in making alloys—with copper forming brass—with tin and copper, bronze, &c. Iron coated with zinc is called galvanized iron.

COMPOUNDS.

1. OF ZINC AND OXYGEN.

ZINC OXIDE.

Formula, ZnO.

Preparation. It occurs as native zinc ore, but is best prepared by the combustion of zinc in presence of oxygen.

Properties. Thus prepared zinc oxide is a white powder, insoluble in water and used chiefly as a pigment. By the addition of one molecule of water, zinc hydrate, $Zn(OH)_2$ is formed. It has strong basic properties, and when acted upon by acids form zinc salts as zinc sulphate ($ZnSO_4$) and zinc carbonate ($ZnCO_3$).

2. ZINC AND CHLORINE.

ZINC CHLORIDE.

Formula, $ZnCl_2$.

Preparation. Zinc chloride is prepared by dissolving zinc in hydrochloric acid (HCl), and evaporating the solution to dryness.

Properties. It is a white crystalline solid,

volatile at high temperatures. It has very strong attractions for water, and hence is a powerful caustic and dehydrating agent.

CADMIUM.

Symbol, Cd. *Atomic weight*, 112. *Equivalence*, II.

Occurrence. It usually occurs in compounds with zinc ores as a sulphide (CdS).

Preparation. As cadmium is more volatile than zinc, it is prepared from the distillation of zinc ores, which contain it by catching the first products of distillation.

Properties. Cadmium closely resembles zinc but is more volatile, is whiter and heavier. It is less easily attacked by acids.

Compounds.

Its compounds also closely resemble those of zinc. The sulphide (CdS) has a bright yellow color, and is used largely as a pigment.

CHAPTER XIII.

CHROMIUM, MANGANESE, ALUMINUM.

CHROMIUM.

Symbol, Cr. *Atomic weight*, 52.5. *Equivalence*, II and VI.

History. This metal received the name of chromium from the beautiful color of most of its compounds.

Occurrence. It occurs in nature combined with oxygen, iron and lead.

Preparation. It is prepared by heating the ore with charcoal, and thus reducing the oxide.

Properties. It is a steel-gray, infusible and extremely hard metal.

Burned in oxygen it forms chromic oxide (Cr_2O_3).

Compounds. As chromium has an equivalence of II or VI it forms *ous* and *ic* compounds by its union with chlorine, oxygen, &c.

But beside one atom of chromium unit-

ing with other atoms to form compounds, two atoms with an equivalence of VI, also unite with the same kind of atoms to form different compounds. Thus there are two chromic oxides, CrO_3 called chromic tri-oxide, and Cr_2O_3 chromic oxide.

CHROMIC TRI-OXIDE.

Formula, CrO_3.

Properties. This oxide exists as crimson, needle-shaped crystals.

It is very deliquiscent and a powerful oxidizing agent. By the addition of a molecule of water chromic acid (H_2CrO_4) is formed.

CHROMIC ACID.

Formula, H_2CrO_4.

Properties. Chromic acid is a pinkish and corrosive fluid.

By its evaporation it loses water, and crystals of chromic tri-oxide CrO_3, are formed.

Chromate salts are formed by replacing the hydrogen in chromic acid by basic elements as sodium and potassium, thus forming sodium chromate (Na_2CrO_4), and potassium chromate K_2CrO_4.

Bi-chromate salts, the most important of which is potassium bi-chromate ($K_2Cr_2O_7$) is formed by replacing the hydrogen in bi-chromic acid ($H_2Cr_2O_7$) with basic elements. This bi-chromic acid is theoretically formed by adding a molecule of water to a double molecule of chromic oxide, $2(CrO_3)$.

CHROMIC OXIDE.

Formula, Cr_2O_3.

Properties. As usually prepared chromic oxide is an amorphous, green powder. It is very hard and insoluble in acids. It is used principally as an indelible ink for coloring bank notes green. Theoretically, there is formed a second chromic acid, ($H_2Cr_2O_4$) by the addition of water to the oxide.

The salts formed from this acid by substituting basic elements for the hydrogen are called chromites. Thus there is ferrous chromite ($Fe\ Cr_2O_4$.)

Chromium sometimes acts as a basic element, and will replace the hydrogen atoms in acids to form salts, as sulphate of chromium ($Cr_2\ (SO_4)_3$.)

MANGANESE.

Symbol, Mn. *Atomic weight*, 55. *Equivalence*, II-IV-VI.

History. Manganese received its name having at first been mistaken for magnetic iron.

Occurrence. It occurs usually in nature as an oxide, but also in combination with sulphur, iron and arsenic.

Preparation. It is prepared by reduction of its oxide (MnO_2) with charcoal at a high temperature.

Properties. It is a grayish white brittle metal, oxidizing readily, forming manganese dioxide (MnO_2).

Compounds. As manganese usually has an equivalence of II or VI, by its union with chlorine, oxygen, &c., are formed ous and ic compounds.

With an equivalence of VI, manganese may unite with other elements either by the use of one or two atoms. Thus manganic trioxide is MnO_3, and manganic oxide is Mn_2O_3.

There is still a third oxide of manganese, which is the one occurring in nature and called manganic dioxide (MnO_2). Manganese in this compound has an equivalence of IV.

Manganic dioxide gives off oxygen readily

upon application of heat, and hence is a good oxidizing agent. Theoretically acids and salts are produced from these oxides in a manner analogous to the formation of similar acids and salts from chromium oxides. But besides the *ate* and *ite* salts thus produced there are also *permanganates* formed by substituting basic elements in place of the hydrogen in *permanganic acid* ($H_2Mn_2O_8$).

The most important of these salts is potassium permanganate ($K_2Mn_2O_8$.) which is a strongly oxidizing agent, and therefore extensively used as a disinfectant.

Manganese acting as a basic element with an equivalence of II will replace the hydrogens in acids, forming salts, as for example sulphate of manganese ($MnSO_4$).

ALUMINUM.

Symbol, Al. *Atomic weight*, 27.5. *Equivalence*, VI.

Occurrence. Aluminum is the most abundant element in nature next to oxygen and silicon with which it is usually found combined. The ruby and sapphire are oxides of aluminum. Clay is a silicate of aluminum.

Preparation. Aluminum is prepared by passing the vapor of aluminum chloride, (Al_2Cl_6) over sodium. Thus:

Alumin. chloride. Sodium Sod. Chloride. Aluminum.

$$Al_2Cl_6 + 6Na = 6NaCl + Al_2.$$

Properties. It is a very light bluish white metal, malleable and ductile. It is not volatile, but will burn at a high temperature forming aluminum oxide (Al_2O_3). It does not tarnish easily.

Uses. Aluminum is made into ornaments and used in forming alloys.

Compounds. Aluminum has strong basic properties and unites by use of two atoms (Al_2), with acids to form salts. The sulphates $Al_2(SO_4)_3$ are the most important compounds, as they are characteristic ingredients of all alums.

An alum is a double sulphate plus water, and may be expressed by the general formula, $Al_2(SO_4)_3 + (x) + H_2O$. The potash alum ($Al_2(SO_4)_3 + K_2SO_4 + 24H_2O$) is the most common. All alums are crystalline and styptic, used principally in dyeing.

CHAPTER XIV.

IRON, COBALT, NICKEL.

IRON (Ferrum).

Symbol, Fe. *Atomic weight*, 56. *Equivalence*, II and VI.

Occurrence. Pure iron occurs very sparingly in nature, but in combination with oxygen (Fe_2O_3 & Fe_3O_4), silicon, sulphur, carbon, &c., it is abundant.

Preparation. Iron is usually prepared by heating to a high temperature alternate layers of iron ore, (ferric oxide, Fe_2O_3), charcoal and limestone. The carbon removes the oxygen. Thus :—

Ferric oxide. Carbon. Iron. Carbon monoxide.

$$Fe_2O_3 + C_3 = Fe_2 + 3CO,$$

while the limestone unites with the silica and other impurities to form a fusible silicate, which collecting above the melted iron can be drawn off in form of "slag."

Properties. Iron thus prepared is run into molds, and hence called "cast" or "pig" iron.

This is not pure iron but contains carbon, manganese, sulphur and other impurities. White pig iron contains more carbon than does the gray, and is therefore whiter in color, harder, more brittle and more fusible than the gray. "Wrought" iron differs from pig iron by being comparatively pure. It is prepared from pig iron by burning out all its impurities by a process of oxidation called "puddling." Wrought iron is bluish gray in color, fibrous in structure and malleable.

"Steel" is a form of iron containing from 2 to 4 per cent. carbon, and is prepared by heating wrought iron to redness with charcoal.

Bessemer steel is a variety of iron containing beside the carbon, manganese. It is prepared by melting spiegeleisen (pig iron containing manganese and carbon) with wrought iron.

Steel resembles wrought iron, but is granular in structure and more malleable and fusible. It also has the peculiar property of becoming very hard and brittle when heated and then suddenly cooled.

Compounds. As iron usually has an equiva-

lence of II or VI, by its union with other elements *ous* and *ic* compounds are formed. In the *ic* compounds, however, iron usually unites with other elements by the use of two atoms ((Fe_2) with an equivalence of VI) instead of one. Thus Ferric carbonate is $Fe_2(CO_3)_3$.

OF OXYGEN AND IRON.

There are three oxides of iron—Ferrous oxide (FeO), ferric oxide (Fe_2O_3), and magnetic oxide (Fe_3O_4).

The ferric oxide is found in large quantities in nature, and the magnetic oxide or loadstone—so named from its magnetic properties, in small quantities.

COBALT.

Symbol, Co. *Atomic weight*, 59. *Equivalence*, II and VI.

Occurrence. It occurs free in meteorites, but is usually found in combination with arsenic.

Preparation. It is prepared by roasting the ore, thus driving off the arsenic which is volatile. The other elements with which it is found are gotten rid of chemically leaving the oxide of cobalt (Co_2O_3) which is

reduced by heating with charcoal, thus leaving the pure cobalt.

Properties. It has a steel-gray color, is magnetic, very hard, oxidizes at a red heat and tarnishes readily in a moist atmosphere.

Compounds. Most of the cobalt compounds have brilliant colors, and hence used as pigments. As cobalt has usually an equivalence of II or VI it forms *ous* and *ic* compounds—the *ous* salts being the more important. In most of the *ic* compounds cobalt unites with other elements by the use of two atoms (Co_2). It has strong basic properties replacing the hydrogens in acids.

OF COBALT AND OXYGEN.

There are three oxides of cobalt analogous to those of iron—cobaltous oxide (CoO), cobaltic tri or sesqueoxide (Co_2O_3,) and cobaltic oxide (Co_3O_4.)

NICKEL.

Symbol, Ni. *Atomic weight,* 59. *Equivalence,* II–IV–VI.

Occurrence. It occurs free in meteoric iron ore, but found usually in combination with arsenic, antimony and sulphur.

Preparation. It is prepared in a manner analogous to the preparation of cobalt.

Properties. Nickel is a silver white metal, ductile and malleable. It does not tarnish in dry air, and is not easily attacked by acids.

Uses. Nickel is used for plating and in making alloys. German silver is an alloy of nickel, copper and zinc.

Compounds. The equivalence, mode of union and compounds of nickel are analogous to those of cobalt. The *ous* salts are more important than the *ic*, the nitrate ($Ni(NO_3)_2$) and sulphate ($NiSO_4$) are the two most important soluble nickel salts.

CHAPTER XV.

PLATINUM, GOLD, SILVER.

PLATINUM.

Symbol, Pt. *Atomic weight*, 197. *Equivalence*, II and IV.

Occurrence. Platinum occurs native in round grains, but is usually impure, containing gold, iron, copper, &c.

Preparation. It is best prepared by melting the impure platinum with lead, which forms a nearly pure alloy of platinic-lead. This alloy is next heated in a current of air. The lead melts, oxidizes and is drawn off as slag leaving the pure platinum as a porous mass.

Properties. Platinum is white, very malleable and ductile metal. A good conductor of heat and electricity and only fusible in an oxy-hydrogen blow pipe flame. It is not attacked by single acids, and is but slowly dissolved in aqua-regia.

It does not tarnish. Finely divided or in the porous form, platinum will absorb 800

times its own volume of oxygen, and thus is an active oxidizing agent.

Compounds. As platinum has an equivalence of II or IV, its compounds are either *ous* or *ic*. Thus there are *two* chlorides of platinum—the platinous chloride ($PtCl_2$) and the platinic chloride ($PtCl_4$), which is formed by dissolving platinum in aqua-regia.

GOLD (Aurum).

Symbol, Au. *Atomic weight*, 196.6. *Equivalence*, I and III.

Occurrence. Gold is found pure in nature in quartz rock or sand.

Preparation. The quartz rock containing the gold is finely pulverized and treated with mercury, which forms an amalgam with the gold. The mercury is distilled off leaving the pure gold.

Properties. Gold is a soft yellow metal, very ductile and malleable. Does not tarnish in the air, and the only acid attacking it is aqua-regia.

Uses. It is used in forming alloys and for coinage and jewelry.

Compounds. Gold having an equivalence

of I or III. forms *ous* and *ic* compounds, the most important of which are the chlorides, AuCl, and $AuCl_3$.

Auric chloride ($AuCl_3$) is produced by dissolving gold in aqua-regia.

The two oxides of gold are aurous oxide, (Au_2O) and Auric oxide (Au_2O_3).

SILVER (Argentum).

Symbol, Ag. *Atomic weight*, 108. *Equivalence*, I and III.

Occurrence. Silver is found free in nature and in combination with sulphur, bromine, chlorine, &c.

Preparation. Silver is often found in combination with lead, and is extracted from it in a manner analogous to that of separating gold from its amalgam with lead. Silver is prepared from its sulphide by roasting the ore with sodium chloride, thus forming chloride of silver. The mass is then ground fine, and agitated with water and scrap iron which reduces the silver to the metallic state. Mercury is then added which takes up the silver and is afterward distilled off leaving the pure metal.

Properties. Silver is a white metal, duc-

tile, malleable, tenaceous and harder than gold. It is the best conductor of heat and electricity known. Does not tarnish in air, but is easily attacked by chlorine, sulphur, nitric acid, &c.

Uses. Silver is used for coinage, and also in making jewelry and alloys.

Compounds. Silver usually acts as a monad. It has strong basic properties replacing the hydrogens in acids to form salts. Thus we have nitrate of silver, $Ag(NO_3)$ sulphate of silver Ag_2SO_4, chloride of silver AgCl, and are chemically all produced by dissolving silver in the corresponding acid, viz: nitric acid (HNO_3), sulphuric acid (H_2SO_4), and hydrochloric acid (HCl).

Many of the salts of silver, especially the nitrate, turn black in presence of sunlight or organic matter, and therefore are used in photography and for hair dyes, indelible inks, &c.

CHAPTER XVI.

LEAD, COPPER, MERCURY.

LEAD (Plumbum).

Symbol Pb. *Atomic weight* 207. *Equivalence*, II and IV.

Occurrence. Lead occurs in nature principally as a sulphide, but also as a carbonate, arsenate, phosphate, &c.

Preparation. It is usually prepared from the sulphide (PbS) by first roasting the ore in presence of oxygen, thus forming an oxide and sulphate of lead. These compounds react upon the remaining unchanged sulphide producing sulphurous oxide (SO_2) and metallic lead. Thus :—

Lead oxide.		Lead sulphide.		Sulph. oxide.		Lead.
$2PbO$	+	PbS	=	SO_2	+	$3Pb$.
Lead sulphate.		Lead sulphide.		Sulph. oxide.		Lead.
$PbSO_4$	+	PbS	=	$2SO_2$	+	$3Pb$.

Properties. Lead is a soft, malleable and bluish white metal. It tarnishes readily in the air and is easily fusible.

Water dissolves lead, especially if it con-

tains organic matter or nitrates, depositing it as a carbonate, $PbCO_3$.

Acetic, carbonic and nitric acids dissolve it readily, but other acids scarcely attack it. It acts as a constitutional poison when taken into the human system.

Uses. Lead is used in forming alloys—with arsenic it forms shot—with antimony and tin, type-metal, &c. It is also used largely in the manufacture of water pipes.

Compounds. Lead usually acts as a basic element with an equivalence of II, and as such replaces the hydrogens in acids to form salts. Among the most important of these salts are the sulphate ($PbSO_3$), the nitrate $Pb(NO_3)_2$, the carbonate ($PbCO_3$) and the acetate ($Pb(C_2H_3O_2)_2$) of lead.

The plumbic acetate is soluble in water, has a sweet astringent taste and is commonly called "sugar of lead." There are two oxides of lead—the plumbic oxide (PbO) or "litharge," which occurs free in nature, and is prepared chemically by melting lead in a current of air; and plumbic dioxide PbO_2. Red lead is a compound of the two oxides ($2PbO + PbO_2$).

Chrome yellow is chromate of lead (Pb-

CrO_4). Many of the lead salts are used as pigments.

COPPER (Cuprum).

Symbol, Cu. *Atomic weight*, 63.5. *Equivalence*, II.

Occurrence. Copper occurs free in nature, but usually in combination with sulphur and iron (pyrites), carbon, oxygen, &c.

Preparation. From its combination with carbon and oxygen metallic copper is prepared by reducing the ores with carbon and silica in a blast furnace. Copper is usually prepared from the pyrites by repeatedly roasting the ore, thus converting a portion of the copper and all of the iron into an oxide. This mass is then fused with silica forming a slag containing the iron, while the copper in the unchanged copper ore remains as a sulphide. By again roasting this sulphide the copper oxide now produced reacts upon the unchanged sulphide, forming metallic copper and sulphurous oxide. **Thus :—**

Copper sulphide. Copper oxide. Sulph. oxide. Copper.

$$Cu_2S + CuO = SO_2 + 3Cu.$$

If any of the oxide remains in the molten

mass it is reduced by "poling"—stirring with a green stick.

Properties. Copper is a reddish colored metal, softer than iron. It is malleable, ductile and tenaceous—a good conductor of heat and electricity. It oxidizes slowly in the air at ordinary temperatures and rapidly upon the application of heat.

It is readily dissolved by nitric acid, and slowly in weak alkaline, acid and saline solutions forming poisonous compounds.—Chlorine and sulphur attack it readily forming chloride and sulphide of copper.

Uses. Copper is used generally in the arts, and for making alloys, as german silver, bronze, &c.

Compounds. Copper is a diad, but unites with other elements to form compounds either with the use of one or two atoms. By the use of one atom (Cu) it forms *ic* compounds, and by the use of two atoms (Cu_2), *ous* compounds. Thus there are two oxides of copper cupric or black oxide (CuO) occurring in nature and always produced by the combustion of copper; and cuprous or red oxide Cu_2O, also occurring free in nature and produced by the reduction of cop-

per solutions by grape sugar in the presence of alkalies. In the same way there are two chlorides of copper, the cupric ($CuCl_2$) and the cuprous (Cu_2Cl_2). Both sulphides of copper (CuS and Cu_2S) occur native.

Ous and *ic* salts of copper are formed by replacing the hydrogens in acids by either Cu_2 or Cu. The *ic* salts are the most stable. Many of the copper compounds are brilliantly colored and used as pigments.

MERCURY (Hydrargyrum).

Symbol Hg. *Atomic weight* 200. *Equivalence*, II.

Occurrence. Mercury occurs free in small quantities, but it is usually combined with sulphur forming cinnabar.

Preparation. Mercury is prepared from cinnabar by roasting the ore, thus volatilizing the metal which is afterward condensed.

Properties. Mercury at ordinary temperatures is a silvery-white fluid metal. It freezes at 40°C, vaporizes at 15°C and boils at 350°C. It does not oxidize at ordinary temperatures. Chlorine, sulphur and nitric acid

(HNO_3) attack it readily, but other acids do not.

Uses. Mercury is used in the manufacture of thermometers and barometers. Its alloys are called amalgams.

Compounds. Mercury like copper is a di-ad and unites with other elements, either by the use of one or two atoms. Its compounds formed by the use of one atom (Hg) are called mercur*ic*, while those formed by the use of two atoms (Hg_2), mercur*ous*.

There are two important chlorides of mercury—the mercuric chloride $HgCl_2$, and the mercurous chloride Hg_2Cl_2. Mercuric chloride Hg=Cl is prepared by heating mercuric sulphate ($HgSO_4$) with sodium chloride (NaCl). Thus :—

Mercuric sulphate. Sod. chloride. Sod. sulphate.

$$HgSO_4 + 2NaCl = Na_2SO_4$$

Mercuric chloride.

$$+ HgCl_2.$$

It is in the form of semi-transparent white crystals soluble in water and very poisonous. It is commonly called bi-chloride of mercury or "corrosive sublimate."

Mercurous chloride, Hg—Cl

Hg—Cl

is prepared by heating mercury with sodium chloride (NaCl), and mercuric sulphate ($HgSO_4$.) Thus :—

Mercuric sulphate. Sod. chloride. Mercury Sod. sulph.

$$HgSO_4 + 2NaCl + Hg = Na_2SO_4$$

Mercurous chlo.

$$+ Hg_2Cl_2.$$

This forms a heavy amorphous powder, insoluble in water and not very poisonous. It is commonly called "calomel."

By heat and exposure to light it decomposes into mercury and mercuric chloride ($HgCl_2$).

There are also two important iodides of mercury, the mercuric iodide HgI_2, and the mercurous iodide Hg_2I_2.

Ous and *ic* salts of mercury are formed by replacing the hydrogens of acids with Hg_2 or Hg.

Thus $Hg(NO_3)_2$ is mercuric nitrate, and $Hg_2(NO_3)_2$ is mercurous nitrate.

The *ic* compounds of mercury are more stable, than the *ous*. "White precipitate," so called, is a mercury amide, $HgClNH_2$ or $(NH_2)—Hg—Cl$.

CHAPTER XVII.

MAGNESIUM, STRONTIUM, BARIUM, CALCIUM.

MAGNESIUM.

Symbol, Mg. *Atomic weight*, 24. *Equivalence*, II.

Occurrence. Magnesium is always found in combination, usually as a hydrate, carbonate, silicate, borate or chloride.

Preparation. It is best prepared by heating magnesium chloride, $MgCl_2$ with sodium. Thus :—

Mag. chloride. Sodium. Magnesium. Sod. chloride.

$$MgCl_2 + 2Na = Mg + 2NaCl.$$

Properties. Magnesium is a silver white metal, malleable and ductile. It oxidizes in moist atmospheres, is volatile and burns with a brilliant white flame at high temperatures.

Compounds. Magnesium has marked basic properties replacing hydrogen in acids to form salts. The oxide of magnesium or "magnesia" (MgO) occurs in nature, and is also

produced by the combustion of magnesium. The sulphate of magnesium or "epsom salts" ($MgSO_4$) is very soluble in water, and is used in medicine. The carbonate of magnesium or "magnesia alba" ($MgCO_3$) is produced by the action of sodium carbonate (Na_2CO_3) upon magnesium sulphate ($MgSO_4$). Thus:—

Mag. sulphate. Sod. carbonate. Mag. carbonate.

$$MgSO_4 + Na_2CO_3 = MgCO_3$$

Sod. sulphate.

$$+ Na_2SO_4.$$

It is a white, light powder, insoluble in water. The hydrate of magnesium $Mg(OH)_2$ occurs free in nature, is slightly soluble in water, and has an alkaline reaction.

STRONTIUM.

Symbol, Sr. *Atomic weight*, 87.5. *Equivalence*, II.

Occurrence. Strontium occurs in nature as a sulphate and carbonate.

Preparation. It is prepared by electrolysis from the chloride of strontium ($SrCl_2$).

Properties. Strontium is a soft yellow metal. It is fusible at a red heat and burns with a brilliant flame forming strontium oxide (SrO).

Compounds, The native salts of strontium, the carbonate ($SrCO_3$), and the sulphate ($SrSO_4$) are insoluble in water and used principally in the preparation of the two soluble salts, the nitrate of strontium ($Sr(NO_3)_2$) and the chloride of strontium ($SrCl_2$).

These salts are used in the arts, and in the preparation of red fire.

BARIUM.

Symbol, Br. *Atomic weight*, 137. *Equivalence*, II.

Occurrence. Barium occurs in nature as a sulphate and carbonate.

Preparation. It is prepared from the chloride of barium ($BaCl_2$) by electrolysis.

Properties. It is a yellowish malleable metal rapidly decomposing water when brought in contact with it.

Compounds. There are two oxides of barium—the barium dioxide (BaO_2) which is used in the preparation of hydrogen dioxide (H_2O_2), and the barium oxide (BaO). The natural salts of barium are the carbonate ($BaCO_3$) and the sulphate ($BaSO_4$) both insoluble in water. The sulphate is used as a paint, and also to adulterate white

lead. Both salts are used in the preparation of the soluble barium salts the nitrate of barium ($Ba(NO_3)_2$) and the chloride ($BaCl_2$).

These compounds are used in making green fire.

CALCIUM.

Symbol, Ca. *Atomic weight*, 40. *Equivalence*, II.

Occurrence. Calcium is one of the most abundant elements in nature. It occurs as a carbonate in chalk, limestone and marble—as a sulphate in gypsum—and as a phosphate, silicate, &c., in bones, minerals and vegetables.

Preparation. It is prepared from the chloride of calcium $CaCl_2$ by electrolysis or by fusing it, or the iodide of calcium (CaI_2) with sodium. Thus:—

Calcium chloride. Sodium. Sod. chloride. Calcium.

$$CaCl_2 + 2Na = 2NaCl + Ca.$$

Properties. Calcium is a brilliant yellow metal, malleable and ductile. It oxidizes and burns brightly at a red heat.

Compounds. One of the principal compounds of calcium is the oxide (CaO) or "lime." It is prepared by driving off car-

bonic oxide (CO_2) from the carbonate of calcium ($CaCO_3$) by means of heat. Thus :—

Calcium carbonate. Calcium oxide. Carbon oxide.

$$CaCO_3 = CaO + CO_2$$

It is a hard, white solid uniting directly with water to form calcic hydrate $Ca(OH)_2$. Calcium hydrate is slightly soluble in water—the solution having an alkaline reaction and commonly called "lime water."

Calcium sulphate ($CaSO_4$) or "gypsum" occurs free in nature. It contains water of crystallization which can be driven off at a temperature of 250°C, forming what is commonly called plaster of Paris or cement.

It is a dry powder, but quickly hardens or sets upon the application of water.

"Mortar" is calcium hydrate $Ca(OH)_2$ and sand. Its hardening properties are chiefly due to its combination with carbonic oxide CO_2 forming calcic carbonate ($CaCO_3$). Thus :—

Calcic hydrate. Carbonic oxide. Calcic carbonate. Water.

$$Ca(OH)_2 + CO_2 = CaCO_3 + H_2O.$$

The water thus produced by this combination evaporates from the walls of new buildings making them very damp for a long time.

Barium, strontium and calcium are called *alkaline earth metals.*

CHAPTER XVIII.

LITHIUM, POTASSIUM, SODIUM, AMMONIUM.

LITHIUM.

Symbol, Li. *Atomic weight*, 7. *Equivalence*, I.

Occurrence. Lithium occurs usually as a chloride in mineral and sea waters. It is also found sparingly in some minerals.

Preparation. It is best prepared from the chloride of lithium LiCl, by electrolysis.

Properties. Lithium is a soft, light and silvery white metal. It tarnishes in the air, decomposes water and burns with a white flame at high temperatures.

Compounds. Many of the lithium compounds are volatile and burn with a brilliant crimson flame. All of its salts are soluble in water, but the carbonate of lithium Li_2CO_3 sparingly so. All of the watery solutions of lithium salts have an alkaline reaction and many are used in medicine.

SODIUM.

Symbol, Na. *Atomic weight*, 23. *Equivalence*, I.

Occurrence. Sodium is a very abundant element in nature. It occurs largely as a chloride in rock salt and sea water. It also occurs as a carbonate, borate, silicate, &c.

Preparation. Sodium is prepared by reducing its oxide (Na_2O) at a white heat with carbon. Thus:

Sodium oxide. Carbon. Sodium. Carb. monoxide.

$$Na_2O + 2C = Na + 2CO.$$

Properties. Sodium is a silvery white soft metal. It decomposes water and oxidizes readily in the air, hence must be kept in some such solution as naptha. By application of heat sodium burns with a yellow flame forming sodium oxide Na_2O.

Uses. Sodium is used in metallurgy as a reducing agent.

Compounds. Sodium chloride or common "salt" (NaCl) is obtained by evaporating saline waters or by mining it in the form of rock salt. Sodium hydrate or "caustic soda" (NaOH) is prepared by the decomposition of water by metallic sodium. Thus:—

Water. Sodium. Sod. hydrate. Hydrogen.

$$H_2O + Na = NaOH + H.$$

or by the action of calcium hydrate $Ca(OH)_2$ upon sodium carbonate Na_2CO_3. Thus :—

Calcic hydrate. Sod. carbonate. Sod. hydrate. Calc. carbonate.

$$Ca(OH)_2 + Na_2CO_3 = 2NaOH + CaCO_3,$$

It is a very deliquiscent white salt, coagulating albumen and having caustic properties. Sodium nitrate or "soda saltpetre" ($NaNO_3$) is a white salt very soluble in water and used in medicine.

Sodium carbonate or "soda" Na_2CO_3 was formerly obtained from the ashes of sea weed, but is now obtained by the famous *Leblanc soda process* as represented in the following formulae. Thus :—

1. Sodium chloride. Sulph. acid. Sod. sulph. Hydrochloric acid.

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl.$$

2. Sodium sulphate. Calcic carb. Carbon. Sodium carbonate.

$$Na_2SO_4 + 2CaCO_3 + 2C = Na_2CO_3$$

Carbonic oxide. Calcic sulphide and Oxide.

$$+ 3CO_2 \qquad + CaS, CaO.$$

"Baking powder is acid, super or bi-carbonate of soda ($NaHCO_3$).

Sodium sulphate (Na_2SO_4). Its preparation is represented in the first formula of Leblanc's soda process.

The acid sodium sulphate ($NaHSO_4$) is prepared by adding sulphuric acid to the

normal salt (Na_2SO_4). All the sodium salts are soluble in water.

POTASSIUM.

Symbol, K. *Atomic weight*, 39. *Equivalence*, I.

Occurrence. Potassium always occurs in combination, usually as a chloride, nitrate or sulphate. It occurs in all vegetable tissues.

Preparation. It is prepared in a manner analogous to that of obtaining sodium, but is a more dangerous process as the metal is apt to combine with carbonic monoxide (CO) which forms an explosive mixture.

Properties. Potassium is a soft bluish white metal, and has properties similar to those of sodium.

Compounds. Potassium chloride (KCl) occurs native and is prepared from sea water. It is much like sodium chloride (NaCl).

Potassium bromide (KBr) and iodide (KI) are prepared by the direct action of bromine and iodine upon potassium hydrate (KOH).

Both salts are used extensively in medicine. Potassium hydrate (KOH) or "caustic potash" is prepared in sticks which are

very deliquiscent and have caustic properties coagulating albumen.

Potassium chlorate ($KClO_3$) is used extensively in the preparation of oxygen.

Potassium nitrate or "nitre" (KNO_3) occurs native upon the surface of soils in Bengal. Potassium sulphate (K_2SO_4) is used extensively in the manufacture of gunpowder ($3C + S + 2KNO_3$).

Potassium carbonate, "potash" or "pearlash" (K_2CO_3) is prepared by evaporating the leachings of wood ashes.

Potassium bicarbonate, called super, acid or bicarbonate of potash, is used extensively in bread making under the name of "saleratus."

All potassium salts are soluble in water and have an alkaline reaction.

AMMONIUM.

Symbol, (NH_4). *Molecular weight*, 18. *Equivalence*, I.

Properties. Ammonium is not an element but is really a compound radical with an equivalence of I. It acts as a basic element with an equivalence of one replacing the hydrogens in acids to form salts.

Ammonium has never been isolated except as an amalgam $(NH_4)_2Hg$, which is prepared, thus :—

Ammon. chloride. Sod, amalg. Sod. chlor. Ammon. amalg.

$$2(NH_4)Cl + Na_2Hg = 2NaCl + (NH_4)Hg.$$

Compounds. Ammonium chloride or "sal ammoniac" $(NH_4)Cl$ is obtained by heating the ammoniacal liquors from coal distillation with hydrochloric acid (HCl).

All the ammonium salts as the carbonate ($(NH_4)_2CO_3$), nitrate ($(NH_4)NO_3$) and sulphate ($(NH_4)_2SO_4$) are soluble in water, alkaline in reaction and volatile.

Sodium, potassium, lithium and the rarer metals caesium and rubidium are generally described together as the *alkaline metals.*

PART III.

EXPERIMENTAL CHEMISTRY.

——:o:——

HYDROGEN.

I. Through the cork of a medium sized glass flask bore a hole and into it fit a bent glass tube from ten to twelve inches long. Put about an ounce of commercial sulphuric acid (H_2SO_4) into the flask, and then slowly add about three times the amount of water. Now add to this solution a small handful of fine pieces of zinc. The following reaction taking place ($H_2SO_4 + Zn = ZnSO_4 + H_2$) the hydrogen is given off immediately, and by fitting the cork in the flask the gas will escape by the delivery tube.

After allowing time for the air in the flask to escape the gas can be ignited, burning with a hot, pale flame; or it may be collected in a receiver, under water, by displacement.

The levity of the gas may be demonstrated by inflating soap bubbles with it.

This is easily done by connecting a common clay pipe with the delivery tube by means of a rubber tube, and then dipping the bowl of the pipe into soap suds. The bubbles will rise rapidly, and if a lighted match is touched to them they will explode—the gas burning with a blue flame.

OXYGEN.

II. Take such a flask as is used in experiment I, and fill about ¼ full of a mixture of powdered chlorate of potash ($KClO_3$) and manganese dioxide (MnO_2) in the proportion of 4 to 1.

By heating this mixture over a sand bath the following decomposition will take place ($KClO_3 = KCl + O_3$), the oxygen being given-off will pass through the delivery tube, and may be collected under water in a receiver by displacement. A piece of charcoal with a spark upon it will burst into a flame when introduced into the oxygen, and dried phosphorus will burn with dazzling brilliancy.

HYDROGEN AND OXYGEN.

III. Hydrogen and oxygen may be made

to unite, as shown in the following experiment :

In experiment I, let the delivery tube pass into a basin of soap suds, previously agitated, so as to form bubbles upon the surface. After more bubbles have formed from the hydrogen gas, touch a match to them, whereupon a loud explosion will follow, denoting the union of the oxygen and hydrogen. Thus : $H_2 + O = H_2O$.

NITROGEN.

IV. Place a small piece of phosphorus upon a cork and float it on a basin of water. Light the phosphorus and cover it with a glass vessel.

The phosphorus will burn the oxygen from the air in the vessel, thus—($P_2 + O_5 = P_2O_5$) forming phosphoric oxide (P_2O_5), while the water will rise in the vessel taking its place.

The white fumes of phosphoric oxide (P_2O_5) will soon disappear, leaving the pure nitrogen. A burning candle placed in this gas will be immediately extinguished.

CARBON.

V. To illustrate the powerful absorption properties of charcoal, float upon a basin of

water a piece of charcoal heated to redness and cover at once with a cylinder. The rise of water in the cylinder will indicate the absorption of the air by the charcoal.

VI. To show the decolorizing property of charcoal, mix a solution of indigo blue with pulverized bone black and filter. The solution will run through the funnel colorless.

CARBON AND OXYGEN.

VII. In the experiment I, add to the acid solution in the flask pieces of marble instead of zinc. Bubbles of carbon-dioxide (CO_2) will be given off, according to the following formula—($H_2SO_4 + CaCO_3 = CaSO_4 + CO_2 + H_2O$). As the gas is much heavier than the atmosphere it can be collected immediately by displacing the air in a vessel.

When thus collected it can be dipped out of the receiver with a cup and then poured upon a lighted candle immediately extinguishing it.

Shaken with lime water ($Ca(OH)_2$) it forms a milky white fluid, calcium carbonate ($CaCO_3$,) being precipitated. Thus:— $Ca(OH)_2 + CO_2 = CaCO_3 + H_2O$.

VIII. To such a flask as used in experiment I, add about an ounce of equal parts of common salt and manganese dioxide (MnO_2) and place over a sand bath.

To this mixture add about two ounces of sulphuric acid (H_2SO_4). The following reaction will take place—($2NaCl + MnO_2 + 2H_2SO_4 = MnSO_4 + Na_2SO_4 + 2H_2O + 2Cl$).

Chlorine will be given off in large quantities through the delivery tube, and being heavier than air may be collected by displacement. A piece of phosphorus introduced into the gas burns brightly as will any of the finely powdered metals as arsenic or antimony.

Dye stuffs, such as prints, moistened and placed in the gas will leave their colors discharged, and writing with common ink will fade rapidly.

IX. To a cylinder one half full of chlorine gas add an equal amount of hydrogen gas. Mix the gases well and apply a lighted match or expose them to strong sunlight. They will unite with a loud explosion, forming hydrochloric acid (HCl).

To prevent accident by breakage, wrap a towel about the cylinder before applying the match.

X. Place in such a flask as used in experiment I, about an ounce of common salt and then two ounces of sulphuric acid (H_2SO_4) and heat gently. The following reaction will take place ($2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl$).

Hydrochloric acid gas (HCl) will be given off rapidly through the delivery tube.

It may be collected by displacement or the gas may be passed through water, forming liquid hydrochloric acid.

IODINE.

XI. To show the reaction of iodine upon starch, add one drop of the tincture of iodine to a weak solution of starch, a deep blue color will be at once developed.

CHLORATES.

XII. The chlorates are powerful oxidizing agents, and will form explosive mixtures with sulphur, sugar, &c.

Take about six grains of finely powdered chlorate of potash ($KClO_3$) and three grains of sulphur, mix well together, wrap in a fine

paper, put on an anvil and strike the mass with a hammer—a loud explosion will follow, or—

XIII. Add to equal parts of chlorate of potash and sugar one drop of sulphuric acid (H_2SO_4).

A loud explosion will follow if we strike the mass a sharp blow with a hammer.

SULPHUR.

XIV. To make sulphuretted hydrogen (H_2S) put about an ounce of iron pyrites (FeS) in such a flask as used in experiment I, and add two ounces of sulphuric acid H_2SO_4.

The following reaction will take place—($FeS + H_2SO_4 = FeSO_4 + H_2S$) the sulphuretted hydrogen being given off through the delivery tube and has an odor of rotten eggs.

If the gas be passed through weak acid solutions of the metals, as arsenic, lead, &c., bright colored precipitates will be formed.

XV. To about an ounce of sulphur add a little alcohol and then set on fire. The following reaction will take place ($S + O_2 = SO_2$) and sulphurous acid (SO_2) be given off.

Its bleaching properties may be illustrated by burning sulphur under a glass shade containing brilliantly colored flowers.

The flowers will gradually lose their color, turning an almost pure white or gray.

The color of many substances bleached with sulphur dioxide may be restored by means of ether or benzol applications.

PHOSPHORUS.

XVI. Place a small fragment of dried phosphorus in the centre of a dinner plate. ignite it by means of a hot wire and cover immediately with a glass vessel.

White fumes of phosphorus oxide (P_2O_5) will fill the jar and particles adhering together will fall to the plate in white flakes resembling snow.

XVII. Melt a small piece of phosphorus in a wine glass of hot water and then pass into it by the delivery tube a current of oxygen as prepared in experiment II. The phosphorus will take fire and burn with a brilliant flame.

XVIII. Dissolve a small piece of phosphorus in carbon disulphide (CS_2) and pour

a little of the solution over perfectly dry filter paper.

Let the solution evaporate rapidly by exposing the paper to heat or by rapidly shaking it. The paper will take fire spontaneously and burn with rapidity.

ARSENIC.

XIX. To the hydrogen generator used in experiment I, add a few drops of a solution of arsenic. After the air is all expelled from the flask light the gas as it is given off from the delivery tube and direct the flame against a piece of cold porcelain.

Metallic arsenic will be deposited upon the surface of the porcelain as a black stain which is soluble in hypo-chlorite of soda (NaClO).

As the gas thus produced (Hydrogen arsenide H_3As) is very poisonous, the experimenter must be careful not to inhale it.

XX. Heat a little arsenous acid with charcoal in the bulb of a glass tube. A metallic mirror of arsenic will form just about the bulb on the sides of the tube.

ANTIMONY.

XXI. Add a solution of antimony to the

hydrogen generator used in experiment I, and proceed as in experiment XIX. A black stain will form on the porcelain, but it is insoluble in the hypochlorites.

SODIUM.

XXII. Throw a small piece of metallic sodium upon the surface of warm water.

It will immediately take fire, burning with a hissing noise and often exploding with a loud report.

The chemical action taking place is as follows—$Na + H_2O = NaOH + H$.

SILVER.

XXIII. To illustrate the readiness with which the soluble chlorides are thrown down by the nitrate of silver ($AgNO_3$), dissolve in a tumbler of water a grain of common salt and add a drop of hydrochloric acid. To this solution add a few drops of nitrate of silver, whereupon a white flocculent precipitate of the chloride of silver ($AgCl$) will immediately form, the reaction taking place as follows—$NaCl + AgNO_3 = AgCl + NaNO_3$.

AMMONIUM.

XXIV. Ammonium hydrate (NH_4OH) is acted upon very readily by some of the iron compounds.

To illustrate, add one drop of aqua ammonia to a tumbler of water and then a few drops of the tincture of the chloride of iron.

A brownish flocculent precipitate will immediately form of ferric hydroxide ($Fe_2(OH)_6$), according to the following reaction which takes place—$6NH_4OH + Fe_2Cl_6 = 6NH_4Cl + Fe_2(OH)_6$.

XXV. Many precipitates will redissolve upon the addition of an excess of one or other of the reagents employed.

Thus for example—aqua ammonia (NH_4OH) added in small quantities to a solution of sulphate of copper ($CuSO_4$) will form a greenish blue precipitate which redissolves readily upon a further addition of ammonia.

ERRATA.

——·o·——

Page 22, line 18, Read nitr*ic*.

" 37, equasion 2, Read

$$H_2SO_4 + Zn = ZnSO_4 + H_2$$

Page 70, last formula. Read

$$(CN)_2 \text{ or } \begin{matrix} C \equiv N \\ | \\ C \equiv N \end{matrix}$$

Page 134, line 7. Read chromic tri-oxide.

INDEX.

—:o:—

A.

Acetylene, 67
Acid molecules, 26
Alkaline earths, 159
" metals, 165
Alkalamine, 30
Allotropic. 79
Alum, 137
Aluminum, 136
" compounds, 137
Amalgams, 152
Ampere's law, 13
Amide, 30
Amine, 30
Ammonia, 53
" amalgam, 165
" type, 30
Ammonium, 164
" compounds, 165
Analysis, 10
Animal economy, 75
Antimony, 121
" oxides of, 122
" acids of, 123
Antimonous oxide, 123

Aqua regia, 58
Arsenic, 115
" acids of, 117
" oxides of, 117
Arsenious acid, 118
" oxide, 117
Arsenic acid, 119
" oxide, 119
Arsenic compounds, 116
Artiads, 18
Artiad radicals, 24
Atoms, 9
Atomic weight, 15
Atomicity, 14
Atmosphere, 51

B.

Baking powder, 162
Barium, 156
" compounds of, 156
Basic molecules, 27
Bismuth, 124
" oxides of, 125
" compounds, 125
Borax, 106
Boron, 104
Boric oxide, 105
" or boracic acid, 105
Boron, compounds of, 105
Bromine, 85
" acids of 86
oxides of, 86

C.

Cadmium, 131
Calcium, 157
" compounds of 157
Cast iron, 138
Calomel, 153
Carbon, 61
" monoxide. 70
" dioxide, 68
" di-sulphide, 94
Carbonator, 74
Carbonic acid, 69
Caustic potash, 163
" soda, 161
Cement, 158
Charcoal, 63
Chemical reactions, 31
Chemistry, 9
Chlorine, 77
" acids of, 81
" oxide of 80
" compounds of, 79
Chlorous oxide, 83
" acid, 83
Chloric acid, 83
Choke damp, 68
Chromium, 132
" oxides of, 133
" acids of, 133
" compounds of, 132
Chromic acid, 133
" oxide, 134
" trioxide, 133

Coal, 63
Cobalt, 140
Combustion, 73
Compound molecules, 10 & 21
Compound radicals, 24
Copper, 149
" compounds of, 150
Corrosive sublimate, 152
Cyanogen, 70

D.

Diamond, 61
Diffusion of gases, 32
Dihydrogen dicarbide, 67

E.

Elements, 10
Elemental molecules, 10
Empirial formula, 21
Epsom salts, 155
Equivalence, 17
Ethylene, 66
Experimental chemistry, 169

F.

Fire damp, 65
Fluorine, 89
Formula, 21
Fowler's solution, 119
Fuming nitric-acid, 58

G.

Glass, 104
Gold, 144
" compounds of, 144

Graphic symbol, 19
Graphite, 62
Green fire, 157
Gun powder, 164
Gypsum, 158

H.

Halogens, 77
Hydrogen, 36
" carbide, 65
" chloride, 79
" bromide, 86
" arsenide, 116
" antimonide, 122
" cyanide, 71
" oxide, 41
" nitride, 53
" dioxide, 47
" di-carbide, 66
" sulphide, 93
" silicide, 102
" phosphide, 109
Hydrocarbons, 64
Hydrochloric acid, 79
" " type, 29
Hydrocyanic acid, 71
Hydrobromic acid, 86
Hydrofluoric " 89
Hydroiodic " 88
Hyponitrous oxide, 56
Hypochlorous " 82
" acid, 82
Hypophosphites, 114

I.

Illuminating gas, 73
Inorganic chemistry, 35
Iodine, 87
" acids of, 88
" oxides of 88
Iodic acid, 88
Iron, 138
" oxides of, 140
" compounds of, 139

L.

Lead, 147
" compounds of, 148
LeBlanc's soda process, 162
Lime, 157
Litharge, 148
Lithium. 160
" compounds of, 160

M.

Magnesia, 154
Magnesium, 154
" compounds of. 154
Manganese. 134
" acids of. 136
" oxides of, 135
" compounds of, 135
Marsh gas, 65
" " type, 30
Mass, 10
Matter, 9
Mercury, 151
" compounds of, 152

Meta-phosphorie acid, 113
" -phosphates, 113
" -silicic acid. 103
Methane, 65
Mineral waters, 42
Molecule, 9
Molecular weight. 15
Mosaic gold, 129
Mortar, 158
Multiplication of atoms, 20

N.

Negative atoms, 25
Nickel, 141
Nitrogen, 50
" acids of, 55
" oxides of, 55
Nitrous acid, 56
Nitric acid, 57
Nitrogen dioxide, 59
" teroxide 59
Nitre, 164

O.

Olifiant gas, 66
Ortho-phosphoric acid, 111
" -phosphates, 111
Oxygen, 38
Ozone, 47

P.

Paris green, 119
Perchloric acid, 84
Perisads, 18
Perisad radicals, 25

Phosphorus, 107
" oxides of, 109
" acids of, 110
Phosphoric oxide, 111
Phosphorus oxide, 113
Phosphine, 109
Pig iron, 138
Plaster of Paris, 158
Platinum, 143
" compounds of, 144
Positive atoms, 25
Potassium. 163
" compounds of, 163
Prussic acid, 71
Pyrophosphoric acid, 112
Pyrophosphates, 112

R.

Radicals, 24
Rational formulae, 21
Red fire, 156

S.

Saleratus, 164
Saline molecules, 27
Salt, 161
Saltpetre, 162
Salts, 28
Saturated molecules. 24
Sheele's green, 119
Silica, 102
Silicon, 101
" oxide of, 102
" acid of, 103
" compounds of, 102

Silver, 145
" compounds of, 146
Simple radicals, 24
Soda, 162
Sodium, 161
" compounds of, 161
Stannous chloride, &c., 127
" oxide, 128
" acid, 128
Stannic chloride, 127
" oxide, 128
" acid, 128
Steel, 139
Strontium, 155
" composed of, 156
Sulphur, 91
Sulphur, oxides of, 94
" acids of, 95
Sulphurous oxide, 95
" acid, 96
Sulphuric oxide, 97
" acid, 98
Sulpho-sulphuric acid, 100
Sulphide of antimony, 124
Sulphuretted hydrogen, 93
Synthesis, 10

T.

Tartar emetic, 128
Tests for arsenic, 119
Theonic acids, 100
Theoretical chemistry, 9
Tin, 126

Tin oxides of, 128
" acids of, &c., 128

U.

Unsaturated molecules, 24

V.

Vegetable economy, 75

W.

Water, 41
" type, 28
White precipitate, 153
Wrought iron, 139

Z.

Zinc, 129
" compounds of, 130